

DOT POINT

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IB CHEMISTRY CORE

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Introduction

What the book includes

This book provides questions and answers for each dot point in the IB Chemistry Core syllabus from the International Baccalaureate Diploma Programme for Chemistry:

- Quantitative Chemistry
- Atomic Structure
- Periodicity
- Bonding
- Energetics
- Kinetics
- Equilibrium
- Acids and Bases
- Oxidation and Reduction
- Organic Chemistry

Format of the book

The book has been formatted in the following way:

1.1 Subtopic from syllabus.

1.1.1 Assessment statement from syllabus.

1.1.1.1 First question for this assessment statement.

1.1.1.2 Second question for this assessment statement.

The number of lines provided for each answer gives an indication of how many marks the question might be worth in an examination. As a rough rule, every two lines of answer might be worth 1 mark.

How to use the book

Completing all questions will provide you with a summary of all the work you need to know from the syllabus. You may have done work in addition to this with your teacher as extension work. Obviously this is not covered, but you may need to know this additional work for your school exams.

When working through the questions, write the answers you have to look up in a different colour to those you know without having to research the work. This will provide you with a quick reference for work needing further revision.

Command Terms and Verbs to Watch

account, account for State reasons for, report on, give an account of, narrate a series of events or transactions.

analyse Interpret data to reach conclusions.

annotate Add brief notes to a diagram or graph.

apply Use an idea, equation, principle, theory or law in a new situation.

assess Make a judgement of value, quality, outcomes, results or size.

calculate Find a numerical answer showing the relevant stages in the working (unless instructed not to do so).

clarify Make clear or plain.

classify Arrange into classes, groups or categories.

comment Give a judgement based on a given statement or result of a calculation.

compare Give an account of similarities and differences between two (or more) items, referring to both (all) of them throughout.

construct Represent or develop in graphical form.

contrast Show how things are different or opposite.

deduce Reach a conclusion from the information given.

define Give the precise meaning of a word, phrase or physical quantity.

demonstrate Show by example.

derive Manipulate a mathematical relationship(s) to give a new equation or relationship.

describe Give a detailed account.

design Produce a plan, simulation or model.

determine Find the only possible answer.

discuss Give an account including, where possible, a range of arguments for and against the relative importance of various factors, or comparisons of alternative hypotheses.

distinguish Give differences between two or more different items.

draw Represent by means of pencil lines.

estimate Find an approximate value for an unknown quantity.

evaluate Assess the implications and limitations.

examine Inquire into.

explain Give a detailed account of causes, reasons or mechanisms.

extract Choose relevant and/or appropriate details.

extrapolate Infer from what is known.

identify Find an answer from a given number of possibilities.

justify Support an argument or conclusion.

label Add labels to a diagram.

list Give a sequence of names or other brief answers with no explanation.

measure Find a value for a quantity.

outline Give a brief account or summary.

predict Give an expected result.

propose Put forward a point of view, idea, argument, suggestion etc for consideration or action.

recall Present remembered ideas, facts or experiences.

show Give the steps in a calculation or derivation.

sketch Represent by means of a graph showing a line and labelled but unscaled axes but with important features (for example, intercept) clearly indicated.

solve Obtain an answer using algebraic and/or numerical methods.

state Give a specific name, value or other brief answer without explanation or calculation.

suggest Propose a hypothesis or other possible answer.

summarise Express concisely the relevant details.

synthesise Put together various elements to make a whole.

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DOT POINT

CORE 1

Quantitative Chemistry

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1.1 The mole concept and Avogadro's constant. © IBO 2007

1.1.1 Apply the mole concept to substances. © IBO 2007

1.1.1.1 The word 'mole' can refer to a little mammal that burrows under ground, a pile or mound of a substance or a small mark on the skin. In chemistry it is used as a unit of measurement.

(a) Recall the chemical definition of a mole.

.....

.....

(b) Outline the relationship between this definition and the Italian scientist, Avogadro.



Amedeo Avogadro (1776-1856)

.....

.....

.....

(c) If you wanted to count out a mole of sand grains, how many would you need to count?

.....

(d) Why do chemists use moles to measure the quantities of chemicals?

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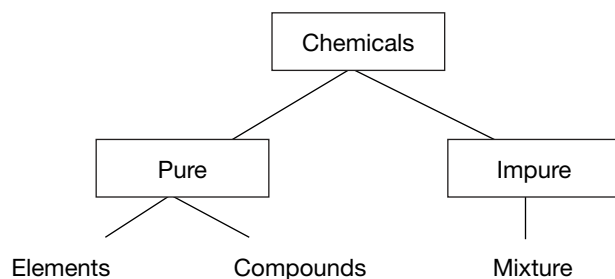
(e) How old would you be if you had lived for 1 mole of seconds? Based on this calculation, would the mole be an appropriate unit for the measurement of time?

.....

.....

.....

- 1.1.1.2** Avogadro's concept of a mole can be applied to all the substances in the Universe. These substances can be classified according to the following scheme.



- (a) Define the types of substances used in this classification and name an example of each.

.....

.....

.....

.....

.....

.....

- (b) Classify each of the following substances, listing them in the correct column of the table below.
Sea water, oxygen, sodium carbonate, magnesium, bread, granite rock, steel, magnesium oxide, hydrogen, iodine, carbon, uranium, aluminium, calcium, sodium, nitrogen, calcium sulfide, blood, carbon dioxide, vinegar, mercury, lemonade, chlorine, dilute hydrochloric acid, iron, margarine.

Pure substance			Impure substance
Element		Compound	Mixture
Metal	Non-metal		

1.1.1.3 All substances in the Universe are made of particles.

(a) Recall the particle theory of matter.

.....

.....

.....

(b) The following diagrams illustrate particles of different types of substances. State whether each of these particles represents an element, mixture or compound. Justify your answers.

(i) 

(ii) 

(iii) 

(iv) 

(v) 

(vi) 

(c) The particles that make up substances can be atoms, ions or molecules. Distinguish between these particles.

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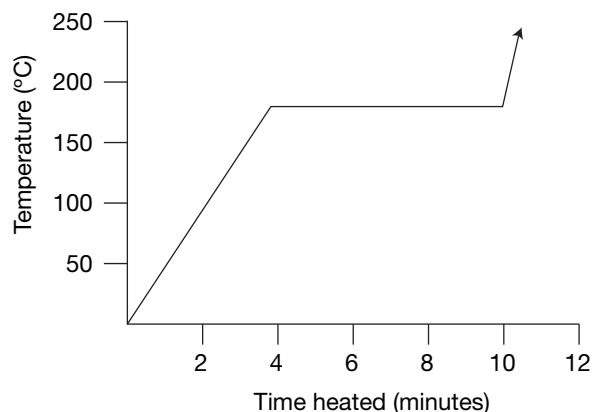
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Answer Questions 1.1.1.4 to 1.1.1.6 by selecting the most correct alternative.

1.1.1.4 The list containing only pure substances is:

- (A) Table salt, unpolluted air, rainwater, lemonade.
- (B) Diamond, copper wire, sodium bicarbonate, oxygen gas.
- (C) Cooking oil, petrol with no additives, cough syrup, sugar crystals.
- (D) Graphite, coal, diamond gemstones, copper ore.

1.1.1.5 The temperature of a white powder was recorded while it was being heated and the results graphed. The powder decomposed at 500°C.



The white powder was a:

- (A) Mixture of elements.
- (B) Mixture of elements and compounds.
- (C) Pure sample of an element.
- (D) Pure sample of a compound.

Revision: Scientific notation and significant figures

In the next section you will be dealing with numerical examples and calculations. Before proceeding, you should check your understanding of scientific notation and significant figures.

1.1.1.6 Which of the following numbers is expressed in correct scientific notation?

- (A) 456×10^8
- (B) 4.56×10^{10}
- (C) 0.456×10^{11}
- (D) 456.00×10^8

1.1.1.7 Write the following numbers in scientific notation.

- (a) 0.0000373
- (b) 57600000
- (c) 659.5

1.1.1.8 Write the following in normal numbers.

- (a) 7.5×10^4
- (b) 2.8×10^{-4}
- (c) 3.7×10^1

1.1.1.9 Identify the number of significant figures in each of the following.

- (a) 0.000025
- (b) 208876
- (c) 0.0208876
- (d) 20.0
- (e) 4×10^3
- (f) 4.0×10^3

1.1.1.10 Perform the following calculations and state your answer with the appropriate degree of accuracy.

- (a) $3.27 + 5.6$
.....
- (b) $6.097 - 4.1768$
.....
- (c) 0.056×0.472
.....
- (d)
$$\frac{500.23}{4.3}$$

.....

1.1.2 Determine the number of particles and the amount of substance (in moles). © IBO 2007

1.1.2.1 Calculate the number of:

- (a) Atoms in 1.00 mole of calcium.
.....
.....
- (b) Atoms in 1.00 mole hydrogen gas.
.....
.....
- (c) Total ions in 1.00 mole of calcium chloride.
.....
.....

(d) Molecules in 3.6 moles of water.

.....
.....

(e) Atoms in 3.6 moles of water.

.....
.....

(f) Outer shell electrons in 1.2 moles of calcium atoms.

.....
.....

1.1.2.2 Calculate the moles of each substance present in the following examples.

(a) 6.02×10^{23} ions

(b) 3.01×10^{23} atoms

(c) 6.02×10^{12} ions

(d) 2.00×10^{23} molecules

(e) 14.27×10^{15} atoms

(f) 10 million pebbles in a stream

(g) 18.000×10^{46} stars

1.2

Formulas. © IBO 2007

1.2.1 Define the terms relative atomic mass (A_r) and relative molecular mass (M_r). © IBO 2007

1.2.1.1

- (a) Define the terms relative atomic mass (
- A_r
-) and relative molecular mass (
- M_r
-).

.....

.....

.....

.....

- (b) Explain why these terms have no associated units.

.....

.....

.....

1.2.1.2

- (a) Find the relative atomic mass for each of the following.

Hint: Use the periodic table at the back of the book.

- (i) Iron
- (ii) Calcium
- (iii) Zinc

- (b) Calculate the relative molecular mass for each of the following.

- (i) Calcium oxide (CaO)
- (ii) Sodium chloride (NaCl)
- (iii) Magnesium chloride (MgCl₂)

- (c) Compare the numerical values for the relative atomic mass and relative molecular mass of oxygen.

.....

.....

.....

Note: If you are having trouble understanding the idea of a mole, try thinking of it as like a loaf of bread. A loaf is a convenient-sized unit for obtaining bread, just as a mole is a convenient unit for measuring chemicals.

Loaf of bread	Mole of chemical
If someone has piled lots of slices of bread on the table and you want to know how many loaves are there, you could find out by calculating: $\frac{\text{Weight of all the slices}}{\text{Weight of 1 loaf}} =$	If you have lots of a chemical and you want to know how many moles are there, you could find out by calculating: $\frac{\text{Weight of all the chemical}}{\text{Weight of 1 mole}} =$

Think about how you can extend this analogy as you learn more about moles.

1.2.2 Calculate the mass of one mole of a species from its formula. © IBO 2007

1.2.2.1 Use the periodic table to determine the mass of 1.0 mole of each of the following substances.

- (a) Iron (Fe).....
- (b) Calcium (Ca)
- (c) Oxygen atoms (O).....
- (d) Oxygen molecules (O₂)
- (e) Calcium oxide (CaO)
- (f) Sodium chloride (NaCl)
- (g) Magnesium chloride (MgCl₂)

1.2.3 Solve problems involving the relationship between the amount of substance in moles, mass and molar mass. © IBO 2007

1.2.3.1

- (a) Define the term molar mass and state the unit for measurement of molar mass.

.....

.....

- (b) Complete the following equation to calculate the number of moles present in a given mass of a substance.

$$\text{Number of moles} = \frac{\text{.....}}{\text{molar mass}}$$

1.2.3.2

- (a) The term molar mass can be used when referring to the atomic mass, molecular mass or formula mass of a substance expressed in grams. Explain when it is appropriate to use each of these terms, giving an example of each.

.....

.....

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.....

.....

- (b) State which of these terms (molar mass, atomic mass, molecular mass and/or formula mass) you could use when referring to the mass of the substances listed below.

(i) 6.02×10^{23} atoms of sodium.

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(ii) 6.02×10^{23} carbon dioxide particles.

.....

(iii) 6.02×10^{23} units of NaCl.

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1.2.3.3 Use the periodic table to determine the number of moles present in each of the following.

Hint: Use $n = m/M$.

(a) 680.0 g nitrogen gas.

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(b) 223 g of carbon dioxide gas.

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(c) 6.0×10^3 g of nitrogen dioxide gas.

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(d) 360.00 g sodium.

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(e) 3264.0 g iron.

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(f) 137.80 g aluminium chloride (AlCl_3).

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1.2.3.4 Calculate the mass of the following amounts of substances.

(a) 0.50 mol sulfur.

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(b) 2.0 mol of copper(II) sulfate (CuSO_4).

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(c) 8.60 mol water (H_2O).

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1.2.3.5

(a) Use the periodic table to determine the molar mass of ammonia (NH_3).

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(b) Identify the number of particles present in 1 mole of ammonia.

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(c) Using the information already obtained, determine the mass of one molecule of ammonia.

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1.2.3.6 Calculate the mass of 1 atom of:

(a) Sulfur.

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(b) Hydrogen.

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Answer Questions 1.2.3.7 to 1.2.3.9 by selecting the most correct alternative.

- 1.2.3.7** Two moles of carbon dioxide (CO_2) at 273 K and 101.3 kPa has:
- (A) A mass of 88.02 g.
 - (B) A volume of 22.4 dm^3 .
 - (C) 6.02×10^{23} molecules.
 - (D) Two CO_2 molecules.
- 1.2.3.8** A 10 g sample of methane gas occupies a volume of 20 dm^3 . What volume would be occupied by 10 g of sulfur dioxide gas at the same temperature and pressure?
- (A) 40.0 dm^3
 - (B) 10.0 dm^3
 - (C) 5.0 dm^3
 - (D) 2.5 dm^3
- 1.2.3.9** Which of the following has the greatest mass?
- (A) 3.00 g of the metal mercury.
 - (B) 3.01×10^{23} atoms of hydrogen gas.
 - (C) 3.00 mol helium gas.
 - (D) 22.4 dm^3 of hydrogen gas at 273 K and 101.3 kPa.

1.2.4 Distinguish between the terms empirical formula and molecular formula. © IBO 2007

1.2.4.1

- (a) Define the terms empirical formula and molecular formula.

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- (b) What do empirical and molecular formulas have in common?

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- (c) How do an empirical and a molecular formula differ?

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1.2.4.2 State whether each of the following is an empirical and/or a molecular formula and in each case, justify your decision.

(a) NaCl

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(b) SiO₂

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(c) H₂O

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(d) CO₂

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(e) C₂H₆

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1.2.4.3 Complete the table by writing empirical formulas for the compounds identified.

Name	Empirical formula	Molecular formula
Carbonic acid		H ₂ CO ₃
Ethene		C ₂ H ₄
Acetylene (ethyne)		C ₂ H ₂
Octane		C ₈ H ₁₈
Hydrogen peroxide		H ₂ O ₂
Methane		CH ₄

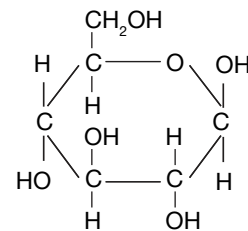
Answer Questions 1.2.4.4 and 1.2.4.5 by selecting the most correct alternative.

1.2.4.4 Which one of the following is an empirical formula?

- (A) H₂O₂
- (B) C₆H₆
- (C) C₃H₈
- (D) N₂O₄

- 1.2.4.5** Glucose is a sugar formed during photosynthesis. The structural formula of glucose is given below. The empirical and molecular formulas for glucose are:

	Empirical formula	Molecular formula
(A)	$C_6H_{12}O_6$	CH_2O
(B)	CH_2O	$C_5H_{10}O_5$
(C)	CH_2O	$C_6H_{12}O_6$
(D)	$(CH_2O)_n$	CH_2O



- 1.2.5 Determine the empirical formula from the percentage composition or from other experimental data.** © IBO 2007

1.2.5.1

- (a) Recall what is meant by the percentage composition of compounds.

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- (b) Calculate the percentage of sodium in table salt (sodium chloride, NaCl).

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- (c) Ammonium sulfate contains 72.69% sulfate ions. Determine the mass of sulfate ions in a 5 gram sample of ammonium sulfate.

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- 1.2.5.2** The compound potassium oxide consists of 83% potassium and 17% oxygen. Find its empirical formula.

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- 1.2.5.3** Sodium hypochlorite is commonly used as a source of pool chlorine. This compound contains 47.62% chlorine and 30.885% sodium. It also contains oxygen. Use this data to determine the empirical formula of sodium hypochlorite.

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- 1.2.5.4** A copper compound in an ore has the percentage composition by weight of copper 51.4%, carbon 9.7% and oxygen 38.9%. Calculate the empirical formula of this compound and state its name.

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- 1.2.5.5** A student burnt 5.00 grams of magnesium in air and found that the mass of the ash produced was 8.29 g.
- (a) Account for the increase in mass from 5.00 g to 8.29 g during this combustion reaction.

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- (b) What mass of oxygen combined with the 5.00 g of magnesium during combustion?

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- (c) Calculate the empirical formula of magnesium oxide.

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- (d) Write an equation for the burning of magnesium in air.

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- 1.2.6 Determine the molecular formula when given both the empirical formula and experimental data.** © IBO 2007

- 1.2.6.1** Two organic compounds, P and Q, each have the same empirical formula of CH_2 . Find the molecular formula for these compounds if their molar masses are found experimentally to be: P = 56.11 and Q = 98.19.

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- 1.2.6.2** The empirical formula of the covalent compound called benzene is CH and the molecular mass of benzene is 78 g. Find the molecular formula of benzene.

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- 1.2.6.3** The percentage composition of the compound ethene is 85.7% carbon and 14.3% hydrogen and its molar mass is 28 grams. Calculate its empirical and molecular formulas.

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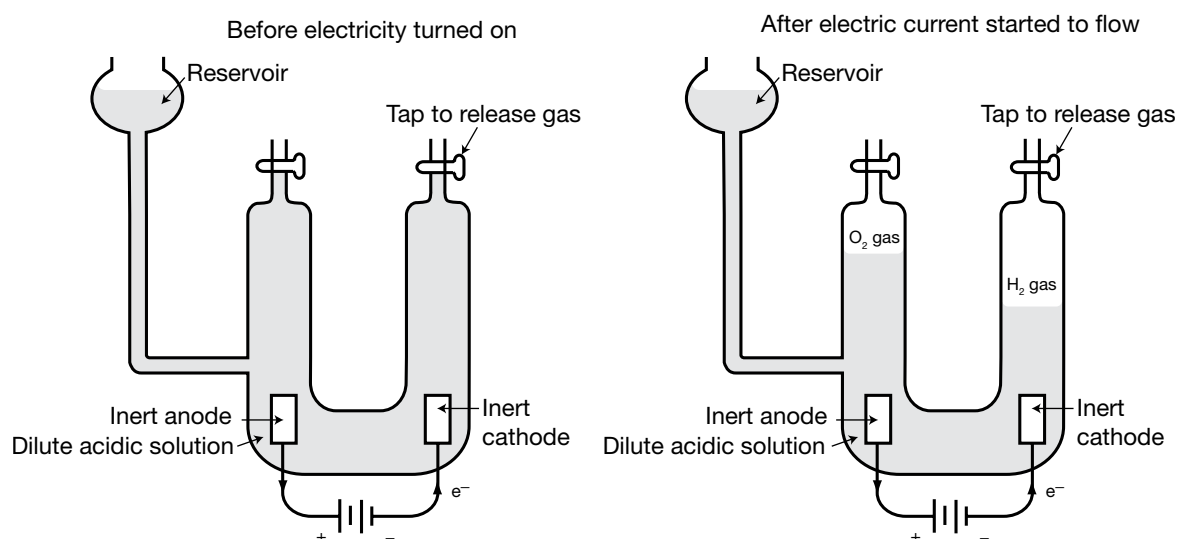
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- 1.2.6.4** A group of students set out to decompose water in order to find its chemical composition. They set up a voltameter, filled it with acidified water and passed an electric current through the water. Gases were produced at each electrode. The students tested these gases and found they were hydrogen and oxygen. Their results are shown in the diagram.



- (a) Name the process the students carried out.
-
- (b) Why was some acid added to the water in the voltameter?
-
- (c) Suggest ways the students could have tested the gases to find out what they were.
-
- (d) What do the results of this experiment suggest about the composition and empirical formula of water?
-

1.2.6.5 Anhydrous copper sulfate is a white powder with the formula CuSO_4 . A group of students placed a weighed sample of blue copper sulfate crystals into a beaker and heated slowly until it became white. When cool, they weighed the chemical again and found that it had lost 36.0% of its mass. They assumed that the drop in mass was all due to the evaporation of water of crystallisation from the copper sulfate crystals.

Using this information, determine the molecular formula of the original hydrated copper sulfate crystals.

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1.3. Chemical equations. © IBO 2007

1.3.1 Deduce chemical equations when all reactants and products are given. © IBO 2007

1.3.1.1 In order to write equations you must first be able to write formulas with correct subscripts.

(a) When writing formulas of compounds it is important to know, or be able to work out, the symbol and valence (combining power) of each element present in a compound. Complete the table to revise this information.

Name of element	Symbol	Valence	Name of element	Symbol	Valence
Sodium		1	Oxygen		
	Ca			S	
Potassium				Sr	
Silver				Cl	
Iodine			Bromine		
	Fe		Fluorine		
	Pb			Al	
Nitrogen				Mg	
Nickel			Barium		
Cobalt				Mn	
Copper				C	

(b) To complete the table, write formulas or names for the common polyatomic ions.

Name of polyatomic ion	Formula	Name of polyatomic ion	Formula
Cyanide		Sulfite	
Phosphate		Nitrate	
Hydrogenphosphate			NO_2^-
	H_2PO_4^-	Chromate	
Carbonate		Dichromate	
	HCO_3^-	Oxalate	
	SO_4^{2-}	Permanganate	
Hydrogensulfate			CH_3COO^-
	OH^-	Ammonium	

1.3.1.2 Write formulas for the following compounds.

- (a) Sodium chloride
- (b) Magnesium sulfate
- (c) Calcium carbonate
- (d) Copper(II) sulfide
- (e) Silver nitride
- (f) Lithium nitrate
- (g) Aluminium bromide
- (h) Hydrochloric acid
- (i) Iron(II) hydroxide
- (j) Barium phosphate

- 1.3.1.3** Complete the table to illustrate the link between the name, formula and composition of the listed compounds.

Name of compound	Formula	Chemical composition
Sulfuric acid		Hydrogen:sulfur:oxygen in ratio 2:1:4
Sodium hydroxide		
Magnesium chloride		
		Hydrogen:oxygen = 2:1
Carbon dioxide		
Chromium(III) sulfite		
	H_2CO_3	

- 1.3.1.4** Write equations in words and symbols to represent the described reactions, indicating the state ((s), (l) (g) or (aq)) for each species.

(a) The synthesis of sodium chloride from sodium metal and chlorine gas.

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(b) The synthesis of liquid water by the combustion of hydrogen gas in the presence of oxygen gas.

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(c) The decomposition of water by electrolysis to form hydrogen gas and oxygen gas.

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(d) The decomposition of green copper(II) carbonate powder, by heating it, to form carbon dioxide and black copper oxide powder.

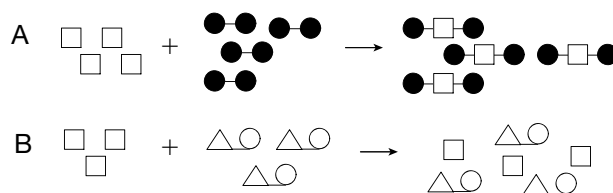
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- (e) When carbon dioxide gas is bubbled through an aqueous solution of calcium hydroxide (limewater), a white precipitate of calcium carbonate is formed as well as water.

- (f) Copper metal reacts with concentrated nitric acid solution to form a deep blue solution of copper(II) nitrate and water. This reaction must be carried out in a fume cupboard because fumes of the toxic brown gas called nitrogen dioxide are released.

1.3.1.5 State which one of the two diagrams below represents a chemical reaction. Justify your answer.



Answer Question 1.3.1.6 by selecting the most correct alternative.

1.3.1.6 Iron(III) oxide reacts with carbon monoxide in a blast furnace producing molten iron and carbon dioxide.

Which alternative gives the correct balanced equation for this reaction?

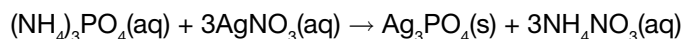
- (A) $\text{Fe}_2\text{O}_3(\text{s}) + \text{CO}(\text{g}) \rightarrow \text{Fe}(\text{s}) + \text{CO}_2(\text{g})$
 (B) $\text{Fe}_2\text{O}_3(\text{s}) + 5\text{CO}_2(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + 5\text{CO}_2(\text{g})$
 (C) $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \rightarrow \text{Fe}_2(\text{s}) + 3\text{CO}(\text{g})$
 (D) $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + 3\text{CO}_2(\text{g})$

1.3.2 Identify the mole ratio of any two species in a chemical reaction. © IBO 2007

1.3.2.1 Balance the following equations by filling in the missing coefficients and, wherever possible, include states of species.

- (a) $\text{K} + \text{H}_2\text{O} \rightarrow \text{H}_2 + \text{KOH}$
 (b) $\text{BaO} + \text{HNO}_3 \rightarrow \text{Ba}(\text{NO}_3)_2 + \text{H}_2\text{O}$
 (c) $\text{Pb}(\text{NO}_3)_2 + \text{KCl} \rightarrow \text{KNO}_3 + \text{PbCl}_2$
 (d) $\text{FeCl}_3 + \text{AgNO}_3 \rightarrow \text{Fe}(\text{NO}_3)_3 + \text{AgCl}$
 (e) $\text{Ba}(\text{NO}_3)_2 + \text{K}_2\text{SO}_4 \rightarrow \text{KNO}_3 + \text{BaSO}_4$
 (f) $\text{Al} + \text{Fe}_2\text{O}_3 \xrightarrow{\text{heat}} \text{Al}_2\text{O}_3 + \text{Fe}$
 (g) $\text{C} + \text{Al}_2\text{O}_3 \xrightarrow{\text{heat}} \text{CO} + \text{Al}$

1.3.2.2 For the balanced equation:



(a) Identify the mole ratios of:

(i) Reactant species.

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(ii) Product species.

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(b) What do we call the balancing numbers in an equation?

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(c) Explain why we balance equations.

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1.3.2.3 For each of the following combustion reactions, state the number of moles of oxygen needed for complete combustion of one mole of the metal.

(a) $4\text{K}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{K}_2\text{O}(\text{s})$

(b) $2\text{Fe}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{FeO}(\text{s})$

(c) $\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{s})$

1.3.2.4 For the reaction, $\text{Ca}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{H}_2(\text{g}) + \text{CaCl}_2(\text{aq})$, assuming complete reaction, how many moles of hydrogen gas would be produced if you added:

(a) 0.5 mol calcium to 1 mole of hydrochloric acid?

(b) 0.2 mol calcium to 0.5 mol hydrochloric acid?

(c) 1 mol calcium to 1 mol hydrochloric acid?

1.3.3 Apply the state symbols (s), (l), (g) and (aq). © IBO 2007

1.3.3.1 Identify the meaning of the following symbols in equations: (s), (l), (g) and (aq).

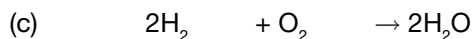
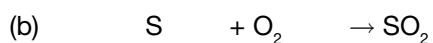
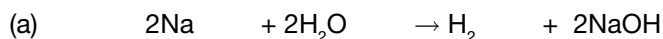
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1.3.3.2 Add symbols to the following equations to show the states of reactants and products.



1.3.3.3 In order to know which substances in a reaction are likely to be solid(s) or in solution (aq), you need to know solubility rules. Identify one solubility rule.

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1.3.3.4 For each of the chemical reactions described below, write a net ionic equation and name any spectator ions.

(a) When calcium metal is placed in a solution of zinc nitrate, a displacement reaction occurs and the products are calcium nitrate solution and particles of solid zinc.

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(b) When pieces of calcium metal are added to dilute sulfuric acid, hydrogen gas is released and a solution of calcium sulfate is obtained.

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(c) Solid calcium carbonate reacts with dilute hydrochloric acid to form carbon dioxide gas, water and dissolved calcium chloride.

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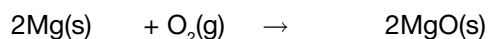
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1.4 Mass and gaseous volume relationships in chemical reactions. © IBO 2007

1.4.1 Calculate theoretical yields from chemical equations. © IBO 2007

1.4.1.1 Magnesium is burnt in air to form magnesium oxide as shown by the following equation.

Magnesium + oxygen $\rightarrow$ magnesium oxide



- (a) Using the periodic table, find the molar mass of magnesium oxide.

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- (b) 2.73 g magnesium undergoes combustion. Convert this mass to moles.

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- (c) Use the equation to determine how many moles of magnesium oxide would be produced.

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- (d) Calculate the mass of magnesium oxide produced.

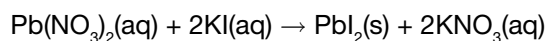
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1.4.1.2 When colourless solutions of lead nitrate and potassium iodide are mixed a yellow precipitate of lead iodide is formed as shown by the equation:



Calculate the mass of lead iodide precipitate formed when 2.63 g of lead nitrate is added to excess potassium iodide.

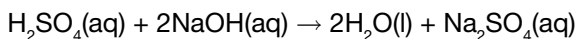
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- 1.4.1.3** Dilute solutions of sulfuric acid and sodium hydroxide react to form sodium sulfate and water as shown by the equation:



- (a) If 0.01 mole of sulfuric acid is used, how many moles of sodium hydroxide will be needed for complete reaction?

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- (b) Calculate the volume of sodium hydroxide solution needed for this reaction if the concentration of the sodium hydroxide solution is 0.5 mol dm^{-3} .

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1.4.2 Determine the limiting reactant and the reactant in excess when quantities of reacting substances are given. © IBO 2007

- 1.4.2.1** Explain what is meant by a limiting reactant.

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- 1.4.2.2** Dilute sulfuric acid reacts with magnesium carbonate with the release of carbon dioxide gas according to the equation:



- (a) According to this equation, 1 mole of magnesium carbonate will react with mole of sulfuric acid to produce mole of carbon dioxide, mole of water and mole of magnesium sulfate.

- (b) Find the molar mass for each of the species in this equation.

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- (c) If the mass of magnesium carbonate used is 7.31 g, how many moles does that represent?

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- (d) If all of this magnesium carbonate is used up in the reaction, how many moles of acid will be needed?

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- (e) 0.01 mol of sulfuric acid is available for this experiment. Will this amount of acid be enough to use up all the magnesium carbonate? Explain.

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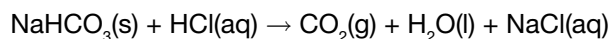
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- (f) What do we call a substance that runs out during a chemical reaction?

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1.4.2.3 Solid sodium hydrogen carbonate is used to neutralise hydrochloric acid as shown by the equation:



- (a) How many moles of sodium hydrogen carbonate will be needed to neutralise 0.04 mol of hydrochloric acid? Explain.

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- (b) Calculate the mass of solid sodium hydrogen carbonate that is needed to neutralise the acid.

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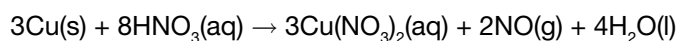
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- (c) If 2.80 g sodium hydrogen carbonate is added to the 0.04 mol of hydrochloric acid, identify the limiting reactant. Explain.

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1.4.2.4 Nitric acid reacts with copper according to the equation:



Identify the reactant which will be left over, and the number of moles left, when you mix each of the following solutions.

- (a) 3.0 mol copper and 7.5 mol nitric acid.

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- (b) 0.25 mol copper and 0.7 mol nitric acid.

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- (c) 0.8 mol copper with 0.8 mol nitric acid.

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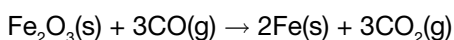
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1.4.3 Solve problems involving theoretical, experimental and percentage yield. © IBO 2007

- 1.4.3.1** An industrial chemist carries out a chemical reaction for which she calculates that the theoretical yield will be 85.00 g of product. On completion, she finds that the actual yield is 80.67 g. Calculate the percentage yield.

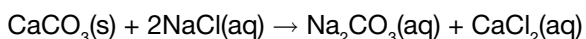
- 1.4.3.2** Hematite is an iron ore containing iron oxide (Fe_2O_3). To extract the iron from this ore it is heated in a blast furnace with carbon monoxide. The reaction can be shown by the equation:



- (a) Calculate the theoretical yield of metallic iron when a total of 1000 tonnes of hematite ore is reacted with excess carbon monoxide.

- (b) The actual yield of iron from the blast furnace is found to be 580 tonnes of iron. Calculate the percentage yield.

- 1.4.3.3** A company in South Australia called Penrice Soda Products, produces 325 000 tonnes per year of soda ash (sodium carbonate). The overall equation for this production process can be written as:



- (a) Assuming 100% efficiency, calculate how many tonnes of calcium carbonate are needed to produce this yield.

- (b) Using the same equation, calculate the weight of calcium chloride that would be produced annually as a by-product.

- (c) If this industrial process was only 78% efficient, calculate the tonnes of calcium carbonate needed to produce the same annual yield.

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1.4.3.4 A sodium hydroxide solution was neutralised with dilute hydrochloric acid, forming sodium chloride and water.

- (a) Write a balanced equation for this reaction.

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- (b) If 0.12 mol of sodium hydroxide reacted completely with 0.12 mol of hydrochloric acid, what would be the mass of salt formed.

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- (c) The resulting solution was evaporated to crystallise the dissolved salt. The recovered salt was then heated, to constant mass, in an oven at 110°C. This ensures the evaporation of any residual water.

The following results were obtained.

Mass of evaporating dish empty = 51.32 g.

Mass of dish + salt after heating = 58.31 g.

Calculate the percentage purity of the salt.

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- (d) Suggest possible reasons why the yield was not 100%.

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1.4.4 Apply Avogadro's law to calculate reacting volumes of gases. © IBO 2007

1.4.4.1 State Avogadro's law.

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1.4.4.2 Three different gases, carbon dioxide, oxygen and ammonia, are collected at the same temperature and pressure. The amount of each gas collected is 0.5 mol. Compare:

(a) The volume (space) they occupy.

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(b) Their mass.

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1.4.4.3 Hydrogen gas burns in oxygen to form liquid water.

(a) Write an equation for this reaction.

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(b) Compare the volumes of reacting hydrogen and oxygen.

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(c) How many litres of oxygen will be needed for complete combustion of 6 litres of hydrogen?

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(d) Based on the equation, can you say that 4 litres of hydrogen will produce 4 litres of water? Justify your answer.

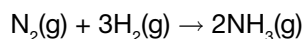
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(e) Compare the moles of hydrogen and oxygen reacting and water produced.

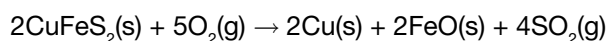
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1.4.4.4 Nitrogen and hydrogen gas react to form ammonia according to the equation:

If 10 cm³ of nitrogen react with 21 cm³ of hydrogen, complete the following table and determine the total volume and composition of the final gas mixture.

Equation	$\text{N}_2(\text{g})$	$+ \quad 3\text{H}_2(\text{g})$	$\rightarrow \quad 2\text{NH}_3(\text{g})$
Reacting volumes	1 vol		
Amount at start		21 cm ³	
Amount used or made			
Amount left at end			

1.4.4.5 The smelting of the copper ore, chalcopyrite can be represented by the following equation:

If 300.00 litres of oxygen is used during this process, calculate:

- (a) The volume of sulfur dioxide produced.

- (b) The mass of sulfur dioxide produced if the gas volumes were measured at STP.

1.4.4.6 Nitrogen oxides are pollutants released into the atmosphere in car exhaust. Nitric oxide (NO) reacts with oxygen to produce nitrogen dioxide.

- (a) Write an equation for this reaction.

- (b) Assuming all volumes are measured at the same temperature and pressure, determine the volume of oxygen needed to react with 200 cm³ of nitric oxide and the volume of nitrogen dioxide formed by this reaction.

1.4.4.7 Methane (CH_4) is the main gas present in natural gas.

(a) Write an equation for the complete combustion of methane in a plentiful supply of oxygen.

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(b) If all gases are measured at the same temperature and pressure, what volume of oxygen will be needed for the combustion of 4 dm^3 of methane?

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(c) Determine the volume of carbon dioxide formed by the combustion of 4 dm^3 of methane.

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1.4.5 Apply the concept of molar volume at standard temperature and pressure in calculations. © IBO 2007

1.4.5.1

(a) Explain what is meant by standard temperature and pressure (STP).

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(b) Define molar volume and identify the value of molar gas volume at standard temperature and pressure.

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.....

(c) Determine the space occupied by each of the following gases at STP.

(i) 1.0 mole of sulfur dioxide.

.....

(ii) 1.0 mole of carbon dioxide.

.....

(iii) 32.0 g of oxygen gas.

.....

(iv) 2.0 moles of nitrogen gas.

.....

(v) 44.01 g of carbon dioxide.

.....

1.4.5.2

- (a) Complete the following equation to calculate the number of moles present in a given volume of a gas at STP.

Number of moles (n) =

- (b) How many moles of each of the following gases are present if the volumes are measured at a temperature of 0°C and atmospheric pressure of 1 atmosphere?

(i) 650 dm^3 of oxygen gas.

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(ii) 650 dm^3 of carbon dioxide gas.

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(iii) $2 \times 10^8\text{ dm}^3$ of sulfur dioxide gas.

.....
.....

- 1.4.5.3** Soft drinks are fizzy because they contain carbon dioxide gas dissolved under pressure. A group of students weighed a bottle of soft drink before it was opened and again after it had been opened and allowed to go flat. They found that the mass of the soft drink decreased by 2.80 g.

- (a) Assuming that all of the decrease in mass was due to escaping carbon dioxide, calculate how many moles of carbon dioxide were lost from the bottle.

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- (b) If the escaping gas was collected and maintained at STP, calculate the volume occupied by this gas.
(*Hint: When performing calculations that carry on from a previous part of the question, always use the full value in your calculator, not the answer you gave for the previous part as this will have been rounded off according to the number of significant figures.*)

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1.4.5.4 Ammonium hydroxide is used in household cleaning products as the ammonium ion is toxic to bacteria. When the bottle is opened some pungent-smelling ammonia gas may be released.

If the ammonia gas released occupies a volume of 50 cm^3 at STP, calculate:

(a) The moles of ammonia released at STP.

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(b) The molar mass of ammonia.

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(c) The mass of this volume of ammonia.

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1.4.5.5 A group of students carry out an experiment designed to collect the oxygen released by the photosynthesis of water plants as shown.

The volume of oxygen produced is measured and found to be 1.3 cm^3 at STP.

(a) Write an equation for the chemical reaction taking place in the plants to produce oxygen.

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(b) Calculate how many moles of oxygen gas would be present in 1.3 cm^3 of the gas at standard temperature and pressure.

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(c) Use the periodic table to determine the molar mass of oxygen gas.

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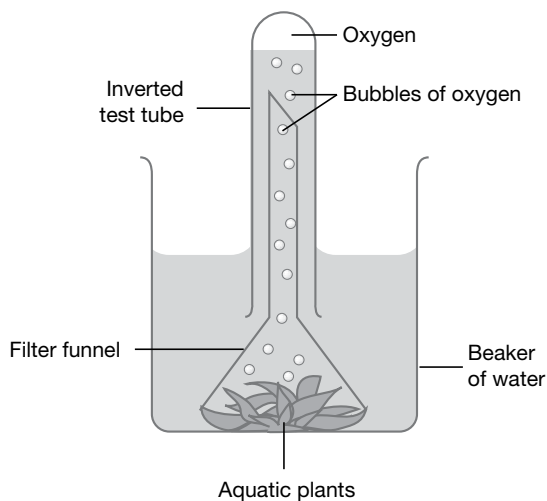
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(d) Calculate the mass of 1.3 cm^3 of oxygen.

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- 1.4.5.6** Carbon dioxide and sulfur dioxide are often released from industrial processes. Complete the following table to convert mass, moles and volumes of these gases at STP.

Name of gas	Mass of gas (g)	Number of moles of gas	Volume of gas at STP (dm ³)
Carbon dioxide	44.01		
Sulfur dioxide	44.01		
Carbon dioxide		5.0 moles	
Sulfur dioxide		5.0 moles	
Carbon dioxide			229.48

- 1.4.6** Solve problems involving the relationship between temperature, pressure and volume for a fixed mass of an ideal gas. © IBO 2007

- 1.4.6.1** (a) Describe the properties of an ideal gas.

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- (b) When are gases least likely to fit the description of an ideal gas? Explain.

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- (c) Which gas would you expect to act more like an ideal gas, carbon dioxide or helium? Justify your answer.

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1.4.6.2 A volume of a gas is enclosed in a rigid container at standard temperature and pressure. Describe and explain the effect on the pressure in the container if:

(a) The number of moles of gas is increased (at the same temperature).

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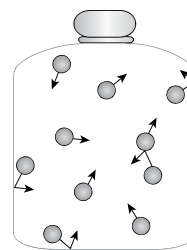
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(b) The temperature is increased.

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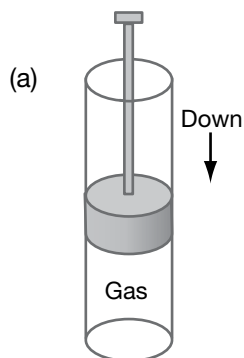
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1.4.6.3 The diagram represents moving molecules of a gas in a closed container with inflexible walls. Add diagrams to represent the movement of molecules inside the container if the conditions are changed as described in the table. Indicate any pressure changes that would occur.

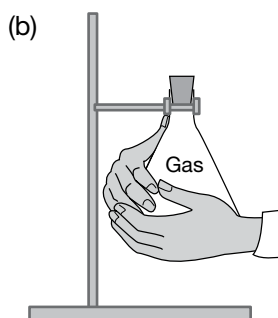


Same container, amount of gas doubled	Same container, temperature increased	Original amount of gas added to container twice as big

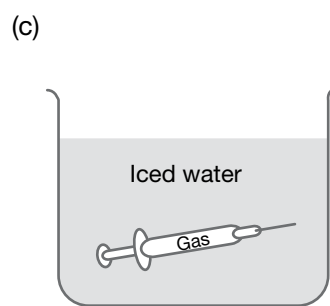
1.4.6.4 For each of the situations illustrated below, indicate whether the pressure of the gas inside the container will increase, decrease or stay the same.



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- 1.4.6.5** For an ideal gas, the relationship between pressure, temperature and volume can be shown by the expression:

$$\frac{PV}{T} = \text{constant} \text{ or } \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

A gas at 300 K and 105 kPa pressure has a volume of 450 cm³. Use the relationship above to calculate the volume of the same mass of gas at 273 K and 100 kPa.

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- 1.4.6.6** A quantity of gas had a volume of 1.3 dm³ at 1 atmosphere pressure and 80°C. Calculate the pressure needed to compress it to 500 cm³ at a temperature of 30°C.

Hint: Remember that temperature must be converted to kelvins.

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1.4.7 Solve problems using the ideal gas equation, $PV = nRT$. © IBO 2007

- 1.4.7.1** The ideal gas law states that the relationship between the amount of gas, its pressure, volume, and temperature is shown by the equation:

$$PV = nRT$$

- (a) State the meaning of the symbols used in this ideal gas law equation and identify their units.

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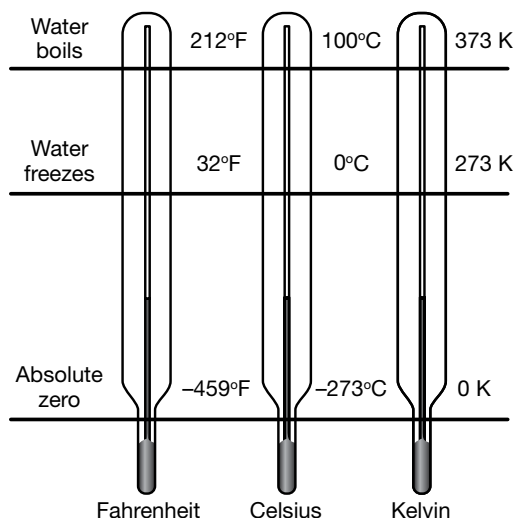
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- (b) When using the ideal gas law equation, the temperature must be expressed in kelvins. Identify the relationship between the numerical values for temperature measured in kelvins and degrees Celsius.

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1.4.7.2 The diagram below shows a comparison of three temperature scales.



The Celsius scale is used in Australia, the Fahrenheit scale in some other countries, e.g. USA, and the Kelvin scale is used mainly by scientists in all countries.

(a) Identify how the positions of 0 and 100 degrees were established on the Celsius scale.

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(b) Explain the concept of absolute zero on the Kelvin scale.

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(c) Research the lowest temperature so far achieved by scientists.

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1.4.7.3

(a) Calculate the number of moles of propane gas in a 40.0 dm³ container at a temperature of 30°C and a pressure of 206 kPa.

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(b) What volume would this amount of propane occupy at STP (0°C and 101.3 kPa).

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1.4.7.4

- (a) What volume would 11.5 g of oxygen occupy at 373 K and 100 kPa?

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- (b) Calculate the volume occupied by the same mass of oxygen as in part (a), kept at the same pressure, if the temperature is decreased to 25°C.

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1.4.7.5

- (a) Calculate the pressure which would be exerted by 250 mol methane, compressed into a 30 dm³ tank at 300 K, assuming methane is an ideal gas.

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- (b) When measured, the actual pressure of the methane was found to be 1.68×10^4 kPa. Compare this measurement with your calculated value and identify two assumptions underlying the ideal gas equation which could account for any difference between calculated and actual pressures.

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- 1.4.7.6** A sample of an unknown gas has a mass of 4.321 g and occupies 24.20 dm³ at STP. Calculate the molecular mass of this gas and identify the gas.

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1.4.8 Analyse graphs relating to the ideal gas equation. © IBO 2007

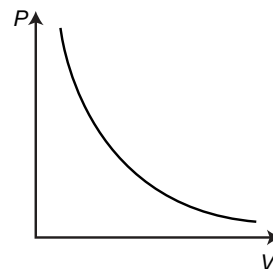
1.4.8.1 The following graphs show relationships between pressure, volume and temperature of a constant mass of an ideal gas. Describe the relationship shown by each graph.

(a) Changes in pressure with changes in volume at a constant temperature.

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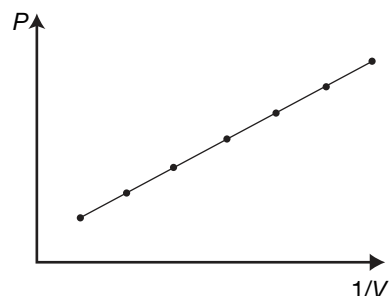


(b) Changes in pressure with changes in 1/volume at a constant temperature.

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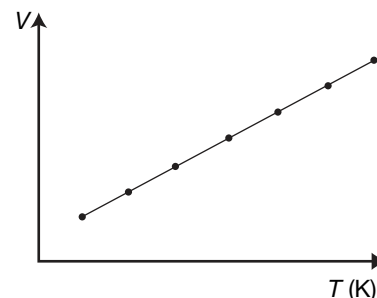


(c) Changes in volume with changes in temperature at a constant pressure.

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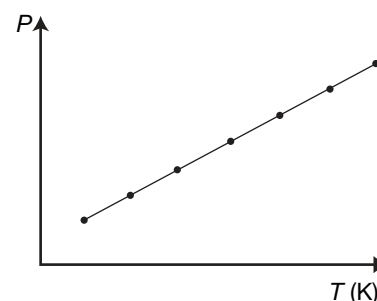


(d) Changes in pressure with changes in temperature with volume kept constant.

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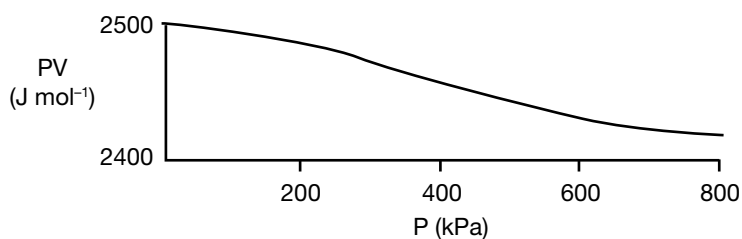
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1.4.8.2 The graph shows the relationship between the pressure of a gas (P) and the product of its pressure and volume (PV).

On the graph draw a line to represent this relationship for an ideal gas.



1.5

Solutions. © IBO 2007

1.5.1 Distinguish between the terms solute, solvent, solution and concentration (g dm^{-3} and mol dm^{-3}).

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1.5.1.1

(a) Define a solution.

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(b) Distinguish between dissolving a pure solid and melting a pure solid.

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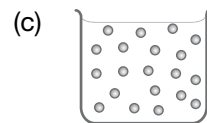
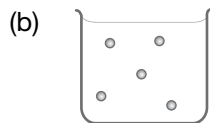
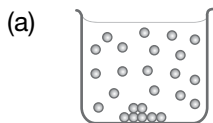
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1.5.1.2 Match the terms with their descriptions in the following table.

Term	Description
(a) Solute	(i) Transparent mixture produced when one substance dissolves in another.
(b) Solvent	(ii) Cannot be dissolved.
(c) Insoluble	(iii) Substance which dissolves.
(d) Soluble	(iv) Can be dissolved.
(e) Solution	(v) Substance which does the dissolving.

1.5.1.3 Label the following diagrams to illustrate the difference between a dilute solution, a concentrated solution and a saturated solution.



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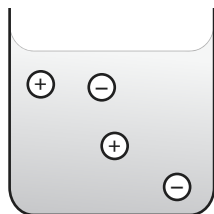
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- 1.5.1.4** Identify four solvents used in the home and an example of a substance that each solvent could dissolve.

Solvent	Substance that could dissolve in the solvent

- 1.5.1.5** Sodium chloride dissolves in water, forming positive sodium ions and negative chloride ions. Two litres of a sodium chloride solution is made up with a concentration of 1 mol dm^{-3} . This is illustrated in diagram A below. Complete diagram B to represent a container of sodium chloride solution with a concentration of 5 mol dm^{-3} .

A 1 mol dm^{-3} NaCl solution



B 5 mol dm^{-3} NaCl solution



- 1.5.1.6** A bottle of household cleaning liquid is labelled as containing 48 milligrams per millilitre of solution. Convert this concentration to g dm^{-3}

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- 1.5.1.7** 2.5 grams of sodium chloride is dissolved in water and the solution is made up to 250 cm^3 .

(a) Calculate the concentration of this solution in g dm^{-3} .

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(b) Use the periodic table to determine the formula mass of sodium chloride.

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- (c) Convert the concentration of the solution from g dm^{-3} to moles dm^{-3} .

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- (d) Explain why the term formula mass has been used rather than molecular mass.

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1.5.1.8 A bottle of mould remover claims a sodium hydroxide concentration of 2.4 g dm^{-3} .

- (a) What mass of sodium hydroxide would be present in a 500 cm^3 bottle of cleaner?

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- (b) Use the periodic table to determine the molar mass of sodium hydroxide.

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- (c) Calculate the concentration of sodium hydroxide in this mould cleaner in mol dm^{-3} .

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1.5.1.9 A bottle of mould remover claims a sodium hypochlorite concentration of 42.0 g dm^{-3} .

- (a) What mass of sodium hypochlorite (NaOCl) would be present in a 1.25 dm^3 bottle of this cleaner?

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- (b) Use the periodic table to determine the molar mass of sodium hypochlorite.

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- (c) Calculate the concentration of sodium hypochlorite in this mould cleaner in mol dm^{-3} .

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1.5.1.10 The number of moles of a dissolved substance present in a solution can be calculated by using the formula:

$$n = c \times V \text{ where}$$

n = number of moles of solute

c = concentration in mol dm^{-3} (also called molarity)

V = volume of solution in dm^3

Use this relationship to calculate the number of moles of solute present in the following solutions.

- (a) 15 dm^3 of solution where the concentration of solute is 0.2 mol dm^{-3} .

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- (b) 5.0 dm^3 of 0.10 mol dm^{-3} sodium chloride solution.

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- (c) 250 cm^3 of 0.60 mol dm^{-3} copper sulfate solution.

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1.5.1.11 Magnesium chloride is dissolved in water to make 100 cm^3 of a solution with a concentration of 0.20 mol dm^{-3} solution.

- (a) How many moles of magnesium chloride would be present in the 100 cm^3 of solution?

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.....

- (b) Use the periodic table to determine the molar mass of magnesium chloride (MgCl_2).

.....

.....

- (c) Calculate the mass of magnesium chloride that was needed to make this solution.

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1.5.1.12 Describe how you would prepare the following solutions.

(a) 500 cm³ of 0.50 mol dm⁻³ Na₂CO₃ solution.

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(b) 100 cm³ of 0.05 mol dm⁻³ hydrochloric acid from 2.00 mol dm⁻³ hydrochloric acid.

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1.5.2 Solve problems involving concentration, amount of solute and volume of solution. © IBO 2007

1.5.2.1 1.73 grams of magnesium reacts with excess dilute hydrochloric acid to produce hydrogen gas.

(a) Write an equation for this reaction.

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(b) Calculate how many moles of magnesium react with the acid.

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(c) Use your equation to determine how many moles of hydrogen will be produced from this many moles of magnesium.

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(d) Determine the mass of the hydrogen produced.

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(e) What volume will this amount of hydrogen occupy at STP?

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1.5.2.2 To determine the accurate concentration of a solution it is often possible to use a volumetric analysis technique called a titration. In this technique, the solution with unknown concentration is reacted with a solution whose concentration is known accurately. Exact volumes of the two reacting solutions are measured at the point where the reaction is complete.

For example, to find the exact concentration of a sodium hydroxide solution, it could be reacted with a measured volume of dilute sulfuric acid with known concentration.

(a) Write an equation for the reaction of dilute solutions of sulfuric acid and sodium hydroxide.

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(b) 50.0 cm³ of 0.20 mol dm⁻³ sulfuric acid is used. How many moles of sulfuric acid does this represent?

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(c) According to the equation in (a), if 50.0 cm³ of 0.20 mol dm⁻³ sulfuric acid is used, how many moles of sodium hydroxide will be needed for complete (stoichiometric) reaction?

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(d) Calculate the concentration of the sodium hydroxide solution if the volume of sodium hydroxide solution needed for this reaction is 25 cm³.

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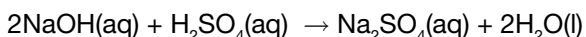
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(e) Suggest a way to determine when this reaction is complete.

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1.5.2.3 Some students wish to determine the concentration of sulfuric acid in a car battery. They extract some of the acid, dilute it accurately by a factor of 5 and carry out four titrations with 0.50 mol dm⁻³ sodium hydroxide solution. The reaction that occurs is shown by the equation below:



The students discard the first titration result then average the results of the next three titrations. They find that, on average, 20 cm³ of the diluted battery acid reacts with 49 cm³ of 0.50 mol dm⁻³ sodium hydroxide solution.

(a) Calculate the concentration of the diluted and the original battery acid.

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- (b) Suggest a reason why the students discarded the first titration result then averaged the results of the next three titrations.

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1.5.2.4 Explain the meaning of the following terms.

- (a) Volumetric analysis.

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- (b) Standard solution.

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- (c) Equivalence point.

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- (d) End point.

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1.5.2.5

- (a) Distinguish between a primary standard and a secondary standard used in a titration.

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- (b) List the requirements for a primary standard.

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(c) Identify the most common primary standards.

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(d) Explain why sodium hydroxide is never prepared as a primary standard for acid-base titrations but instead is used as a secondary standard.

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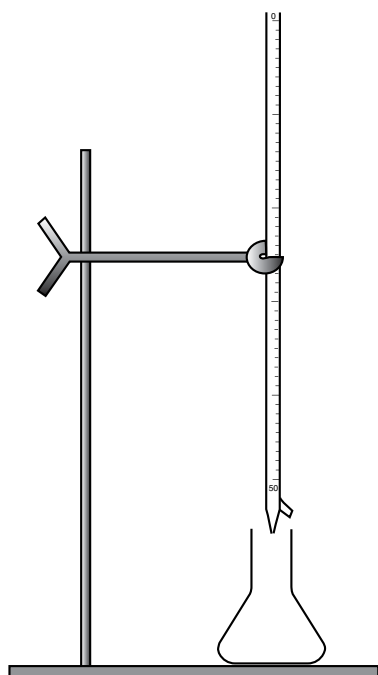
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1.5.2.6 Complete the following table to identify a suitable indicator to use with each of the following acid-base combinations and justify your choice.

Acid	Base	Suitable indicator	Justification
HCl	KOH		
H ₂ SO ₄	NH ₄ OH		
H ₂ CO ₃	NaOH		
CH ₃ COOH	Ca(OH) ₂		

1.5.2.7 The diagrams show equipment you probably used when carrying out acid-base titrations. Identify each diagram and label diagram (a).

(a)



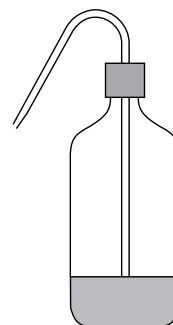
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(b)



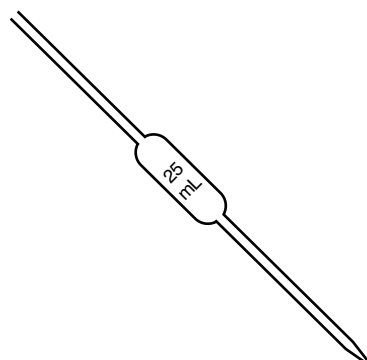
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(c)



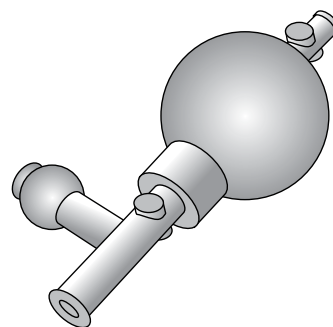
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(d)



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(e)



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1.5.2.8 Complete the following table to summarise the use of a pipette and a burette.

Factor	Pipette	Burette
Function		
Procedure for washing before a titration		
Name given to measured volume		

1.5.2.9

- (a) During an acid-base titration, the acid and base are usually mixed in a conical flask. This flask is rinsed with water only. It is not rinsed with the solution to be placed in it. Explain.

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- (b) Explain why it is not correct procedure to blow or shake the last drop of liquid from a pipette.

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- (c) Identify three other ways to minimise experimental error.

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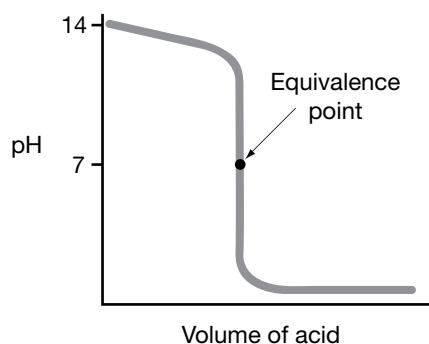
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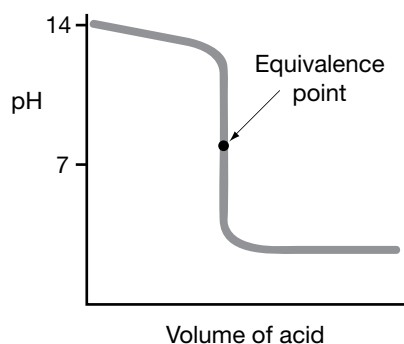
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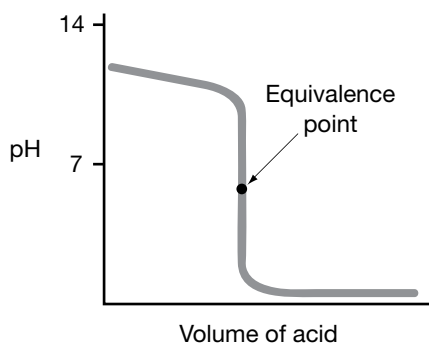
1.5.2.10 The following titration curves represent titrations of strong or weak acids with strong or weak bases. Indicate which type of titration is illustrated by each graph.



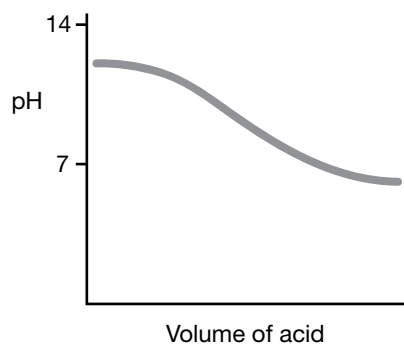
(a)



(b)



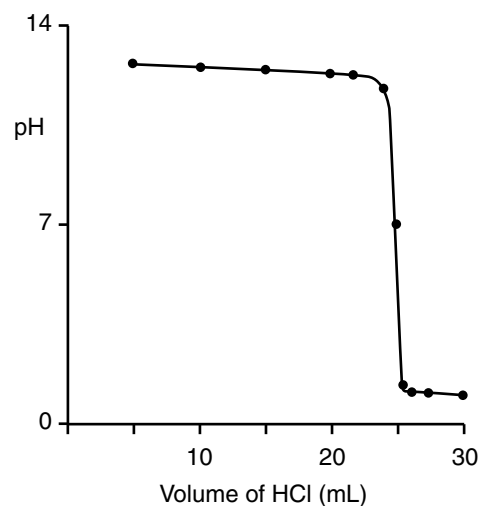
(c)



(d)

1.5.2.11 Using computer technology, a group of students perform a titration using 25 mL 0.1 mol dm⁻³ sodium hydroxide against a solution of hydrochloric acid and they print out the following titration curve. Account for the shape of this graph.

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1.5.2.12 Explain why, when carrying out a titration using potassium permanganate solution, an indicator does not need to be added.

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1.5.2.13 Write relevant equations and calculate the volume of 0.15 mol dm^{-3} sodium hydroxide that will just neutralise:

(a) 25.0 cm^3 of 0.06 mol dm^{-3} hydrochloric acid.

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(b) 25.0 cm^3 of 0.06 mol dm^{-3} sulfuric acid.

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1.5.2.14 A standard solution of anhydrous sodium carbonate was prepared by dissolving 3.25 g in distilled water and then making it up to 250 cm^3 in a volumetric flask.

(a) Calculate the:

(i) Molar mass of anhydrous sodium carbonate (Na_2CO_3).

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(ii) Number of moles in 3.25 g of anhydrous sodium carbonate.

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(iii) Concentration (in mol dm^{-3}) of the sodium carbonate solution.

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- (b) This standard sodium carbonate solution was then used to standardise hydrochloric acid. A pipette was used to measure 25.0 cm^3 of the sodium carbonate solution. This was placed in a conical flask with methyl orange as indicator. The sodium carbonate solution was neutralised by 11.8 cm^3 of the hydrochloric acid. Calculate the following, showing all working.

(i) Moles of sodium carbonate used in the titration.

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(ii) Moles of HCl reacting.

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(iii) Concentration of the hydrochloric acid.

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1.5.2.15 A standard solution of anhydrous sodium carbonate was prepared by dissolving 12.9 g in distilled water and then making it up to 250 cm^3 in a volumetric flask.

(a) Calculate the concentration of the sodium carbonate solution.

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(b) This standard sodium carbonate solution was then used to standardise hydrochloric acid. A pipette was used to measure 25 cm^3 of the sodium carbonate solution. This was placed in a conical flask with methyl orange as indicator. The sodium carbonate solution was neutralised by 24.90 cm^3 of the hydrochloric acid. Calculate the concentration of the hydrochloric acid.

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1.5.2.16 A scientist wishes to analyse the iron content of an ore. All of the ore is present as Fe^{2+} ions dissolved in acid solution. The scientist titrates this solution against a solution of 0.1 mol dm^{-3} potassium permanganate solution.

(a) Write the equation for this reaction.

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(b) If 27.71 cm^3 of the potassium permanganate solution was used in the titration, calculate the number of moles of:

(i) MnO_4^- ions used.

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(ii) Fe^{2+} ions present in the solution being analysed.

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(c) Calculate the mass of iron present in the titrated solution.

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(d) If the solution titrated was derived from a sample of ore with a mass of 1.20 grams, calculate the percentage of iron in the original sample of iron ore.

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Just for fun

1. Use the key provided to solve the puzzle below.

Key

a	b	c	d	e	f	g	h	i	j	k	l	m
Ÿ	∞	≈	≠	∅	Δ	Π	ℓ	Σ	△	ƒ	K	≡
n	o	p	q	r	s	t	u	v	w	x	y	z
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Puzzle

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Solution

2. The answer to each clue below is a nine-letter word. Each answer can be made by combining three of the three-letter groups in the grid. Each three-letter group is only used once.

Three-letter groups

TOR	EMP	CTA	NDS	SYN
IRI	SCR	SIS	SUB	IPT
POU	SPE	THE	COM	CAL

Clues

- (a) A formula which tells us the simplest ratio of atoms or ions in a compound.

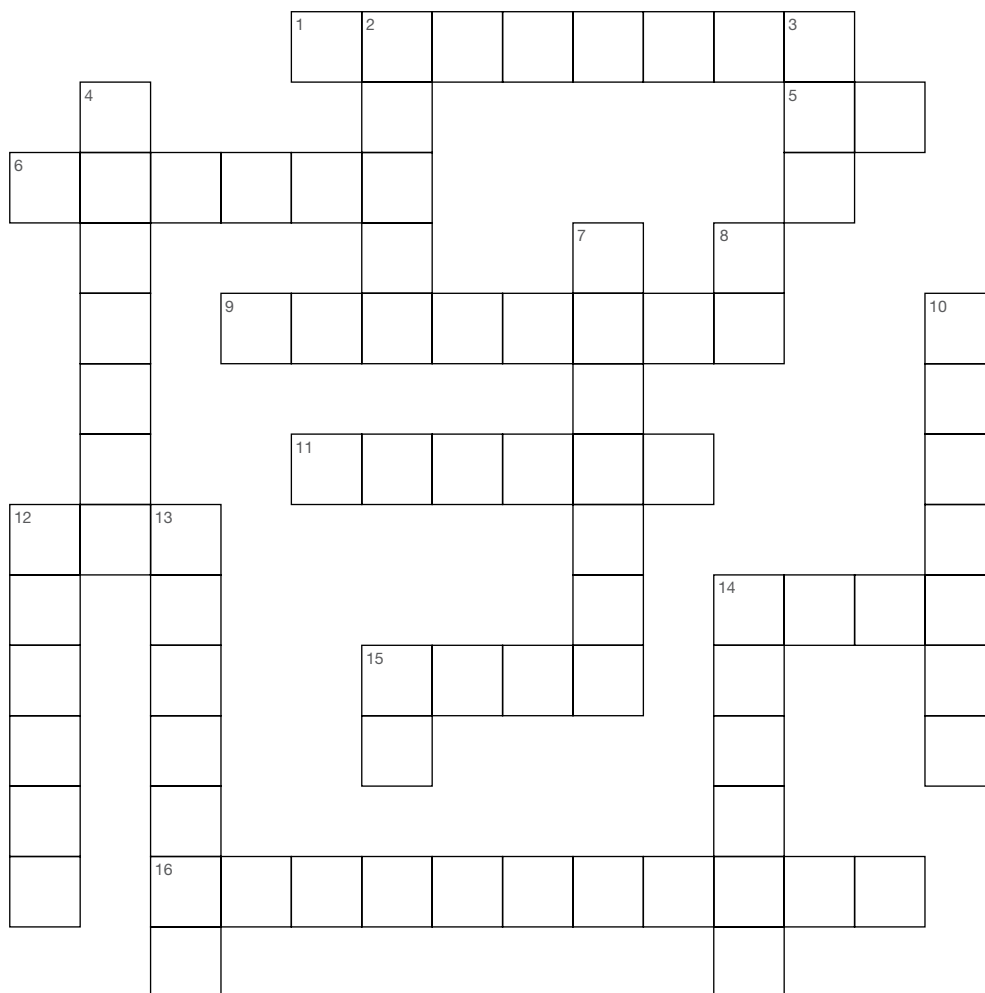
- (b) Manufacture of a compound from simpler substances.

- (c) Substances which contain two or more types of atoms combined in a fixed ratio.

- (d) Ions which are present but not involved in a chemical reaction.

- (e) In the equation $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, what do we call the small number after each H?

3. Crossword.



Across

1. Describes a reaction that is completely used up in a chemical reaction.
5. The symbol for silver.
6. The space occupied.
9. Cl is the symbol for this element.
11. This gets dissolved.
12. Abbreviation for a temperature of 0°C and 101.3 kPa.
14. The amount of a substance that contains Avogadro's number of particles.
15. This word describes substances which are elements or compounds.
16. The number in front of the formula for a species in a balanced chemical equation.

Down

2. A gas which obeys the law expressed as $PV = nRT$.
3. This can diffuse and be compressed.
4. It does the dissolving.
7. Contains two or more types of particles not chemically combined.
8. Symbol for the gas most similar to an ideal gas.
10. Contains only one type of atom.
12. The symbol for this element is Na.
13. Substance made by a chemical reaction.
14. Everything in the Universe.
15. Symbol for lead.

DOT POINT

Answers



This image shows a full page of white paper with horizontal blue or grey ruling lines. The lines are evenly spaced and run across the width of the page, providing a template for handwriting practice. There are no margins, text, or other markings on the page.

CORE 1 Quantitative Chemistry

- 1.1.1.1**
- (a) A mole is the amount of a substance with a mass equal to its formula mass in grams.
 - (b) One mole of a substance contains 6.02×10^{23} atoms, ions or molecules. This number is called Avogadro's constant after the Italian scientist Amedeo Avogadro who determined this number.
 - (c) 6.02×10^{23} sand grains.
 - (d) Chemists study atoms, ions and molecules. These are very tiny and you need a vast number of them to make a quantity large enough to see with the naked eye so we use a mole.
 - (e) Ignoring the existence of leap years, and assuming a year is 365 days, you would live to be 1.910×10^{16} years old. This is longer than the Earth has existed. The mole would not be an appropriate unit for measuring time.

- 1.1.1.2**
- (a) A mixture (impure substance) contains two or more types of particles, which are not chemically combined. Its composition, and thus properties, can vary, e.g. a solution of salt in water.

A pure substance (element or compound) contains only one type of particle.

An element contains only one type of atom, e.g. a sample of pure sodium contains only sodium atoms. An element cannot be changed to any simpler substance by a chemical reaction. Each element has its own symbol, e.g. iron, Fe. An element can be a metal, e.g. sodium, iron and magnesium, or a non-metal, e.g. oxygen, carbon and sulfur.

A compound contains two or more types of atoms which are always combined in the same fixed ratio, e.g. water – H_2O – contains hydrogen and oxygen atoms chemically joined, always in the ratio 2:1.

(b)

Pure substance		Impure substance
Element		Mixture
Metal	Non-metal	
Magnesium	Oxygen	Sea water
Uranium	Hydrogen	Blood
Aluminium	Iodine	Granite rock
Calcium	Carbon	Bread
Sodium	Nitrogen	Steel (an alloy)
Mercury	Chlorine	Vinegar
Iron		Lemonade
		Dilute hydrochloric acid
		Margarine

- 1.1.1.3**
- (a) Every substance (all matter) is made of tiny particles which are constantly moving. These particles may be atoms, ions or molecules.
 - (b) (i), (iii) and (iv) are compounds – each consists of two or more different particles (atoms) joined together in a fixed ratio.
(ii) is an element – two identical particles (atoms) are joined representing a diatomic molecule of an element.
(v) is an element – two identical particles, not joined, representing monatomic molecules.
(vi) is a mixture – there are two different types of particles (atoms) and they are not joined together.
 - (c) An atom is the smallest unit of an element that can take part in a chemical reaction, e.g. an atom of calcium.

An ion is a charged particle formed when an atom gains or loses electrons. When an atom gains one or more electrons it forms a negatively charged ion, e.g. the chloride ion (Cl^-). When an atom loses one or more electrons it forms a positively charged ion, e.g. the magnesium ion (Mg^{2+}).

A molecule is a particle that can exist and move independently. A molecule is made of one or more atoms held together by covalent bonds (shared electrons). Monatomic molecules consist of 1 atom e.g. helium and argon (He and Ar); diatomic molecules consist of 2 atoms, e.g. oxygen and hydrogen (O_2 and H_2); polyatomic molecules contain more than 2 atoms, e.g. water and ammonia (H_2O and NH_3).

1.1.1.4 B

1.1.1.5 D

1.1.1.6 B

- 1.1.1.7**
- (a) 3.73×10^{-5}
 - (b) 5.76×10^7
 - (c) 6.595×10^2
- 1.1.1.8**
- (a) 75 000
 - (b) 0.00028
 - (c) 37
- 1.1.1.9**
- (a) 2
 - (b) 6
 - (c) 6
 - (d) 3
 - (e) 1
 - (f) 2
- 1.1.1.10**
- (a) 8.9 (*Note: When adding, your answer should have no more numbers after the decimal place than any number in the question. Don't forget to round up if the number after the last reportable digit is 5 or more.*)
 - (b) 1.920 (*Note: When subtracting, your answer should have no more numbers after the decimal place than any number in the question. The number after the last reportable digit was less than 5 so it is dropped.*)
 - (c) 0.026 or 2.6×10^{-2} (*Note: Multiplying: 0.056 only has 2 significant figures, so the answer can only have 2 significant figures.*)
 - (d) 1.2×10^2 (*Note: Dividing: 4.3 only has 2 significant figures, so the answer must have the same number.*)
- 1.1.2.1**
- (a) 6.02×10^{23} atoms
 - (b) $2 \times 6.02 \times 10^{23} = 12.04 \times 10^{23}$ atoms (there are 2 atoms in each molecule)
 - (c) $3 \times 6.02 \times 10^{23} = 18.06 \times 10^{23}$ ions ($\text{CaCl}_2 \rightarrow \text{Ca}^{2+} + 2\text{Cl}^-$)
 - (d) $3.6 \times 6.02 \times 10^{23} = 2.2 \times 10^{24}$ molecules
 - (e) $3.6 \times 3 \times 6.02 \times 10^{23} = 6.5 \times 10^{25}$ atoms (there are 3 atoms in each molecule of H_2O)
 - (f) $2 \times 1.2 \times 6.02 \times 10^{23} = 1.4 \times 10^{24}$ electrons (2 outer shell electrons in each calcium atom)
- 1.1.2.2**
- (a) 1.00 mole of ions
 - (b) 0.500 moles of atoms
 - (c) 1.00×10^{-11} moles of ions
 - (d) 3.32×10^{-1} moles of molecules
 - (e) 2.37×10^{-8} moles of atoms
 - (f) 1.7×10^{-17} moles of pebbles
 - (g) 2.99×10^{23} moles of stars
- 1.2.1.1**
- (a) Relative atomic mass (A_r) – the ratio of the average mass per atom of an element compared to 1/12 of the mass of an atom of the carbon-12 isotope.
Relative molecular mass (M_r) – the ratio of the average mass per molecule compared to 1/12 of the mass of an atom of the carbon-12 isotope. It is calculated by adding the atomic masses of the elements in a molecular formula.
 - (b) These values are a comparison (with atoms of carbon-12), not a measured amount. Ratios or comparisons have no units.
- 1.2.1.2**
- (a)
 - (i) 55.85
 - (ii) 40.08
 - (iii) 65.37
 - (b)
 - (i) 56.08
 - (ii) 58.44
 - (iii) 95.21
 - (c) Relative atomic mass = 16, whereas the relative molecular mass = 32. This is because oxygen is diatomic – each molecule of oxygen gas consists of two oxygen atoms – O_2 .

- 1.2.2.1**
- (a) 55.85 g mol⁻¹
 - (b) 40.08 g mol⁻¹
 - (c) 16.00 g mol⁻¹
 - (d) 32.00 g mol⁻¹ (each oxygen molecule is diatomic)
 - (e) 56.08 g mol⁻¹
 - (f) 58.44 g mol⁻¹
 - (g) 95.21 g mol⁻¹
- 1.2.3.1**
- (a) Molar mass is the mass in grams of one mole of a molecule, or the formula mass of an ionic compound. The molar mass of a substance contains 6.02×10^{23} particles. The unit for molar mass is g mol⁻¹.
 - (b) Number of moles = $\frac{\text{mass}}{\text{molar mass}}$ or $n = \frac{m}{M}$
 where n = number of moles, m = mass and M = molar mass.
- 1.2.3.2**
- (a) Molar mass is a general term. It refers to the mass of 1 mole of any substance, whether it is an element, ionic compound or covalent substance.
 Atomic mass is used when referring to the mass of the atoms making up one mole of an element, e.g. carbon = 12.01 g, helium = 4.00 g, magnesium = 24.31 g.
 Molecular mass is used when referring to a mole of molecules of a covalent substance, e.g. oxygen gas = 32 g, ammonia = 17.04 g, concentrated sulfuric acid = 98.08 g.
 Formula mass is used when referring to a mole of a substance with a lattice structure. This may be an ionic substance, e.g. sodium chloride = 58.44 g, or a metallic substance, e.g. Na = 22.99 g. (Note that formula mass is not used for covalent substances.)
 (Note: Each of the examples given above can also be correctly referred to as a molar mass.)
 - (b) (i) Molar mass and atomic mass.
 (ii) Molar mass and molecular mass.
 (iii) Molar mass and formula mass.
- 1.2.3.3**
- (a) 24.27 mol
 - (b) 5.07 mol
 - (c) 1.3×10^2 mol
 - (d) 15.65 mol
 - (e) 58.4 mol
 - (f) 1.03 mol
- 1.2.3.4**
- (a) 16 g
 - (b) 3.2×10^2 g
 - (c) 1.55×10^3 g
- 1.2.3.5**
- (a) 17.04 g
 - (b) 6.02×10^{23}
 - (c) 6.02×10^{23} molecules of ammonia have a mass of 17.04 g
 So 1 molecule = 2.83×10^{-23} g
- 1.2.3.6**
- (a) 5.33×10^{-23} g
 - (b) 1.68×10^{-24} g
- 1.2.3.7** A
- 1.2.3.8** C
- 1.2.3.9** C
- 1.2.4.1**
- (a) Empirical formula – the simplest ratio of atoms or ions in a compound, e.g. CH₂.
 Molecular formula – formula showing the actual number of atoms in a molecule of a molecular covalent compound, e.g. C₂H₄
 - (b) Both empirical and molecular formulas tell us which elements are present in a compound.

- (c) An empirical formula tells us the simplest whole number ratio of elements present in a compound.

The molecular formula tells us the actual number of atoms of these elements in one molecule. The molecular formula is a multiple of the empirical formula.

- 1.2.4.2**
- (a) NaCl – empirical – represents the ratio of sodium and chloride ions as 1:1.
Ionic substances do not exist as molecules (they form a lattice of repeating units) so they do not have a molecular formula.
- (b) SiO_2 – empirical – a giant macromolecule, so only has an empirical formula.
- (c) H_2O – both – the ratio of atoms in the molecule is as simple as possible (empirical formula) and it is also the actual number of atoms in one molecule (molecular formula).
- (d) CO_2 – both – the ratio of atoms in the molecule is as simple as possible (empirical formula) and it is also the actual number of atoms in one molecule (molecular formula).
- (e) C_2H_6 – molecular formula – the empirical formula would be CH_3 .

1.2.4.3

Name	Empirical formula	Molecular formula
Carbonic acid	H_2CO_3	H_2CO_3
Ethene	CH_2	C_2H_4
Acetylene (ethyne)	CH	C_2H_2
Octane	C_4H_9	C_8H_{18}
Hydrogen peroxide	HO	H_2O_2
Methane	CH_4	CH_4

1.2.4.4 C

1.2.4.5 C

- 1.2.5.1**
- (a) Percentage composition tells us the percentage of each element present in a compound. It provides a relative measure of the masses of each different element present in the compound.
- (b) Na in $\text{NaCl} = 22.99/58.44 \times 100 = 39.34\%$
- (c) $72.69/100 \times 5 = 4 \text{ g}$ (Note: 5 g only has 1 significant figure.)

1.2.5.2 The ratio of K atoms to O atoms (percentage/molar mass)

$$= 83/39.1:17/16.0$$

$$= 2.12:1.06$$

(Simplify by dividing both numbers by the smaller number)

$$= 2.12/1.06:1.06/1.06 = 2:1$$

Thus the formula is K_2O .

1.2.5.3 $\text{Na}:\text{O}:\text{Cl} = 30.885/22.99:21.495/16.00:47.62/35.45$

$$= 1.3:1.3:1.3$$

$$= 1:1:1$$

So the formula is NaOCl .

1.2.5.4 CuCO_3 , copper(II) carbonate.

- 1.2.5.5**
- (a) When the magnesium burned, it combined with oxygen from the air to form magnesium oxide. The extra mass is the oxygen that combined with the magnesium during combustion.
- (b) $8.29 - 5.00 = 3.29 \text{ g}$
- (c) The ratio of Mg atoms to O atoms = $5.00/24.31:3.29/16.0$
 $= 0.21:0.21 = 1:1$
 So the formula of magnesium oxide is MgO .
- (d) $2\text{Mg(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{MgO(s)}$

1.2.6.1 Molar mass $\text{CH}_2 = 14.03 \text{ g}$

P $n(\text{CH}_2) = 56.11$

$n \times 14.03 = 56.104$

$n = 4$, so molecular formula is C_4H_8

Q $n(\text{CH}_2) = 98.19$

$n \times 14.03 = 98.19$

$n = 7$, so molecular formula is C_7H_{14}

1.2.6.2 C_6H_6

1.2.6.3 Empirical CH_2 , molecular formula C_2H_4 .

- 1.2.6.4**
- (a) Electrolysis (– to bring about the decomposition of water).
 - (b) Pure water is a poor conductor. With the addition of some acid it becomes an electrolyte – it is a better conductor of electricity than pure water. The acid provides ions to carry the charge.
 - (c) Place lit match in each gas – in oxygen it will burn more brightly, in hydrogen it will go ‘pop’ (explosive reaction), other gases, e.g. carbon dioxide and chlorine, would put the match out.
 - (d) The volume of hydrogen = $2 \times$ the volume of oxygen. So water is made of hydrogen and oxygen in the ratio 2:1 approximately and its empirical formula is H_2O .

1.2.6.5 Molar mass $\text{CuSO}_4 = 159.61$

Let X = the mass of water of crystallisation

$36/100 \times (159.61 + X) = X$

$57.4596 + 0.36X = X$

$0.64X = 57.4596$

$X = 89.78 \text{ g}$

Molar mass of water = 18.02 g

$n = 89.78/18.02 = 4.98$, thus $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ (copper sulfate pentahydrate)

1.3.1.1

(a)

Name of element	Symbol	Valence	Name of element	Symbol	Valence
Sodium	Na	1	Oxygen	O	2
Calcium	Ca	2	Sulfur	S	2
Potassium	K	1	Strontium	Sr	2
Silver	Ag	1	Chlorine	Cl	1
Iodine	I	1	Bromine	Br	1
Iron	Fe	2 or 3	Fluorine	F	1
Lead	Pb	2	Aluminium	Al	3
Nitrogen	N	3	Magnesium	Mg	2
Nickel	Ni	2	Barium	Ba	2
Cobalt	Co	2	Manganese	Mn	2
Copper	Cu	1 or 2	Carbon	C	4

(b)

Name of polyatomic ion	Formula	Name of polyatomic ion	Formula
Cyanide	CN^-	Sulfite	SO_3^{2-}
Phosphate	PO_4^{3-}	Nitrate	NO_3^-
Hydrogenphosphate	HPO_4^{2-}	Nitrite	NO_2^-
Dihydrogenphosphate	H_2PO_4^-	Chromate	CrO_4^{2-}
Carbonate	CO_3^{2-}	Dichromate	$\text{Cr}_2\text{O}_7^{2-}$
Hydrogencarbonate	HCO_3^-	Oxalate	$\text{C}_2\text{O}_4^{2-}$
Sulfate	SO_4^{2-}	Permanganate	MnO_4^-
Hydrogensulfate	HSO_4^-	Ethanoate	CH_3COO^-
Hydroxide	OH^-	Ammonium	NH_4^+

- 1.3.1.2**
- (a) NaCl
 - (b) MgSO_4
 - (c) CaCO_3
 - (d) CuS
 - (e) Ag_3N
 - (f) LiNO_3
 - (g) AlBr_3
 - (h) HCl
 - (i) $\text{Fe}(\text{OH})_2$
 - (j) $\text{Ba}_3(\text{PO}_4)_2$

1.3.1.3

Name of compound	Formula	Chemical composition
Sulfuric acid	H_2SO_4	Hydrogen:sulfur:oxygen in ratio 2:1:4
Sodium hydroxide	NaOH	Sodium:oxygen:hydrogen = 1:1:1
Magnesium chloride	MgCl_2	Magnesium:chlorine = 1:2
Water	H_2O	Hydrogen:oxygen = 2:1
Carbon dioxide	CO_2	Carbon:oxygen = 1:2
Chromium(III) sulfite	$\text{Cr}_2(\text{SO}_3)_3$	Chromium:sulfur:oxygen = 2:3:9
Carbonic acid	H_2CO_3	Hydrogen:carbon:oxygen = 2:1:3

- 1.3.1.4**
- (a) Sodium + chlorine → sodium chloride
 $2\text{Na}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{NaCl}(\text{s})$
 - (b) Hydrogen + oxygen → water
 $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$
 - (c) Water → hydrogen + oxygen
 $2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$
 - (d) Copper carbonate → copper oxide + carbon dioxide
 $\text{CuCO}_3(\text{s}) \rightarrow \text{CuO}(\text{s}) + \text{CO}_2(\text{g})$
 - (e) Calcium hydroxide + carbon dioxide → calcium carbonate(s) + water
 $\text{Ca}(\text{OH})_2(\text{aq}) + \text{CO}_2(\text{g}) \rightarrow \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$
 - (f) Copper + nitric acid(aq) → copper nitrate + water + nitrogen dioxide
 $\text{Cu}(\text{s}) + 2\text{HNO}_3(\text{aq}) \rightarrow \text{CuNO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{NO}_2(\text{g})$

1.3.1.5 A represents a chemical reaction because it shows a new substance being formed by rearranging the atoms of the two original substances. Mole ratio of reactants is 1:1.

B shows a physical (not chemical) change – the two different substances are just being mixed together.

1.3.1.6 D

- 1.3.2.1**
- (a) $2\text{K}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2(\text{g}) + 2\text{KOH}(\text{aq})$
 - (b) $\text{BaO}(\text{s}) + 2\text{HNO}_3(\text{aq}) \rightarrow \text{Ba}(\text{NO}_3)_2(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - (c) $\text{Pb}(\text{NO}_3)_2(\text{aq}) + 2\text{KCl}(\text{aq}) \rightarrow 2\text{KNO}_3(\text{aq}) + \text{PbCl}_2(\text{s})$
 - (d) $\text{FeCl}_3(\text{aq}) + 3\text{AgNO}_3(\text{aq}) \rightarrow \text{Fe}(\text{NO}_3)_3(\text{aq}) + 3\text{AgCl}(\text{s})$
 - (e) $\text{Ba}(\text{NO}_3)_2(\text{aq}) + \text{K}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{KNO}_3(\text{aq}) + \text{BaSO}_4(\text{s})$
 - (f) $2\text{Al}(\text{s}) + \text{Fe}_2\text{O}_3(\text{s}) \xrightarrow{\text{heat}} \text{Al}_2\text{O}_3(\text{s}) + 2\text{Fe}(\text{s})$
 - (g) $3\text{C}(\text{s}) + \text{Al}_2\text{O}_3(\text{s}) \xrightarrow{\text{heat}} 3\text{CO}(\text{g}) + 2\text{Al}(\text{s})$

- 1.3.2.2**
- (a) (i) Ammonium phosphate:silver nitrate in the ratio 1:3.
 (ii) Silver phosphate:ammonium nitrate in the ratio 1:3.
 - (b) Coefficients of the reaction.

- (c) A balanced equation shows us the reaction stoichiometry – ratio of the moles of substances that react and are produced. Balancing an equation is based on the law of conservation of matter, which tells us that matter is neither made nor destroyed during a chemical reaction. So if there are, for example, 3 moles of silver atoms present before a reaction, then there must still be 3 moles of silver atoms present after the reaction.
- 1.3.2.3** (a) 0.25 mol
(b) 0.5 mol
(c) 1 mol
- 1.3.2.4** (a) 0.5 mol
(b) 0.2 mol
(c) 0.5 mol
- 1.3.3.1** The symbols refer to the state of the reactant or product: (s) – solid; (l) – liquid (molten state); (g) – gas; and (aq) – dissolved in water.
- 1.3.3.2** (a) $2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow \text{H}_2\text{(g)} + 2\text{NaOH(aq)}$
(b) $\text{S(s)} + \text{O}_2\text{(g)} \rightarrow \text{SO}_2\text{(s)}$
(c) $2\text{H}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{H}_2\text{O(l)}$
(d) $\text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{H}_2\text{O(l)} + \text{NaCl(aq)}$
- 1.3.3.3** Various, e.g.
Salts of sodium, potassium, ammonium are all soluble.
Nitrates are all soluble.
Chlorides are usually soluble, except AgCl and PbCl₂.
Sulfates are usually soluble, except PbSO₄, CaSO₄, SrSO₄ and BaSO₄.
- 1.3.3.4** (a) $\text{Ca(s)} + \text{Zn}^{2+}\text{(aq)} \rightarrow \text{Ca}^{2+}\text{(aq)} + \text{Zn(s)}$
The spectator ions are nitrate ions.
(b) $\text{Ca(s)} + 2\text{H}^+\text{(aq)} \rightarrow \text{H}_2\text{(g)} + \text{Ca}^{2+}\text{(aq)}$
The spectator ions are sulfate ions.
(c) $\text{CaCO}_3\text{(s)} + 2\text{H}^+\text{(aq)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)} + \text{Ca}^{2+}\text{(aq)}$
The spectator ions are chloride ions.
- 1.4.1.1** (a) Molar mass MgO = 24.31 + 16.00 = 40.31 g
(b) Moles of Mg = 2.73/24.31 = 0.11 mol Mg
(c) 0.11 mol MgO
(d) Molar mass MgO = 40.31 g
Mass of MgO = 0.112299466 × 40.31 = 4.53 g MgO
- 1.4.1.2** 3.66 g
- 1.4.1.3** (a) 0.02 mol NaOH
(b) 40 mL
- 1.4.2.1** If reactants are not present in stoichiometric quantities, there will be some chemicals left over. The chemical which is all used up is called the limiting reactant or reagent because, as soon as one reactant is all used up, the reaction stops. The limiting reactant is the chemical which limits the amount of reaction that can take place.
- 1.4.2.2** (a) 1,1,1,1
(b) $\text{MgCO}_3 = 84.32 \text{ g}$
 $\text{H}_2\text{SO}_4 = 98.086 \text{ g}$
 $\text{CO}_2 = 44.01 \text{ g}$
 $\text{H}_2\text{O} = 18.016 \text{ g}$
 $\text{MgSO}_4 = 120.38 \text{ g}$
(c) No. of moles = mass/molar mass
= 7.31/84.32 = 0.086693548 mol MgCO₃
= 8.67×10^{-2} mol MgCO₃

- (d) 1 mol magnesium carbonate needs 1 mol sulfuric acid (from the equation)
 So 0.086693548 mol MgCO_3 will need 0.086693548 mol H_2SO_4
 $= 8.67 \times 10^{-2}$ mol H_2SO_4
- (e) No, this will not be enough acid. 0.0867 mol of H_2SO_4 is needed, and there is only 0.01 mol available, so all of the MgCO_3 will not be used up. They will run out of sulfuric acid.
- (f) A limiting reactant/reagent – there is not enough of it, so it limits the reaction. H_2SO_4 is the limiting reactant/reagent in this reaction.

1.4.2.3

- (a) 0.04 mol
 (From the equation 1 mol HCl can be neutralised by 1 mol NaHCO_3 so 0.04 mol HCl will be neutralised by 0.04 mol NaHCO_3)
- (b) Molar mass of $\text{NaHCO}_3 = 84.008$ g
 0.04 mol NaHCO_3 will have a mass $= 84.008 \times 0.04 = 3.4$ g
- (c) Sodium hydrogen carbonate will be the limiting reactant as only 2.80 g is provided and 3.36 g is needed for complete reaction.

1.4.2.4

- (a) $0.1875 \text{ mol Cu} = 1.9 \times 10^{-1} \text{ mol Cu}$
- (b) $0.033 \text{ mol} = 3.3 \times 10^{-2} \text{ mol HNO}_3$
- (c) 0.5 mol Cu

1.4.3.1

Percentage yield = actual yield/theoretical yield $\times 100\%$
 $= 80.67/85.00 \times 100 = 94.91\%$

1.4.3.2

- (a) 1 mol iron oxide $\rightarrow$ 2 mol iron
 $(2 \times 55.85 + 3 \times 16.00) \text{ g iron oxide} \rightarrow 2 \times 55.85 \text{ g iron}$
 $159.70 \text{ g Fe}_2\text{O}_3 \rightarrow 111.7 \text{ g Fe}$
 $1000 \text{ tonnes Fe}_2\text{O}_3 \rightarrow 1000 \times 111.7/159.70 \text{ tonnes Fe}$
 Theoretical yield = 699.4 tonnes Fe
- (b) % yield $= 580/699.44 \times 100 = 83\%$ yield (to 2 significant figures).

1.4.3.3

- (a) $\text{CaCO}_3(\text{s}) + 2\text{NaCl}(\text{aq}) \rightarrow \text{Na}_2\text{CO}_3(\text{aq}) + \text{CaCl}_2(\text{aq})$
 Calculate moles:
 From the periodic table:
 Molar mass of $\text{Na}_2\text{CO}_3 = 2 \times 22.99 + 12.01 + 3 \times 16.00 = 105.99$ g
 No. of mol in 325 000 tonnes $= 325\,000 \times 1\,000\,000/105.99 = 3\,066\,327\,000$ mol
 From the equation:
 1 mol Na_2CO_3 produced from 1 mol CaCO_3
 3 066 327 000 mol Na_2CO_3 produced from 3 066 327 000 mol CaCO_3
 Moles to mass:
 $1 \text{ mol CaCO}_3 = 40.08 + 12.01 + 3 \times 16.00 = 100.09$ g
 $3\,066\,327\,000 \text{ mol CaCO}_3 = 100.09 \times 3\,066\,327\,000 = 3.0690866 \times 10^{11} \text{ g}$
 $= 3.07 \times 10^5 \text{ tonnes CaCO}_3$
- (b) Molar mass of $\text{CaCl}_2 = 110.98$ g
 From the equation:
 1 mol CaCO_3 produces 1 mol CaCl_2
 Mass $\text{CaCl}_2 = 325\,000 \times 110.98/105.99 = 3.40 \times 10^5 \text{ tonnes CaCl}_2$ annually.
- (c) 100% efficiency needs 3.07×10^5 tonnes CaCO_3
 78% efficiency would need $3.07 \times 10^5 \times 100/78$
 $= 3.94 \times 10^5 \text{ tonnes CaCO}_3$ annually.

- 1.4.3.4**
- (a) $\text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$
- (b) $0.12 \text{ mol NaCl} = 0.12 \times (22.99 + 35.45) = 7.0128 \text{ g}$
 $= 7.0 \text{ g}$ (to 2 significant figures)
- (c) Theoretical yield = 7.0128 g
 The actual mass of pure salt produced = $58.31 - 51.32 = 6.99 \text{ g}$
 $\% \text{ purity of salt} = 6.99/7.0128 \times 100 = 99.67\%$
- (d) Various, e.g.
 Some salt may have been lost when evaporating the salt solution by spitting.
 Some salt may have been left behind when transferring solutions to different containers.
 There may have been impurities in the NaOH or HCl used.
 Errors of measurement.

1.4.4.1 Avogadro's law states that there is a direct relationship between the volume of a gas and the number of particles present ($V \propto n$). At a specified temperature and pressure, equal volumes of (ideal) gases contain equal numbers of particles. This also means that equal amounts (moles) of gases will occupy the same volume, if they are at the same temperature and pressure.

- 1.4.4.2**
- (a) They will all have the same volume (occupy the same space) as they are all gases at the same temperature and pressure.
- (b) Their masses will be different. $0.5 \text{ mol CO}_2 = 22.01 \text{ g}$; $0.5 \text{ mol O}_2 = 16.00 \text{ g}$; $0.5 \text{ mol NH}_3 = 8.52 \text{ g}$.

- 1.4.4.3**
- (a) $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O(l)}$
- (b) The reacting volume of hydrogen will be twice the reacting volume of oxygen.
- (c) 3 litres.
- (d) No. Water is in the liquid state, it is not a gas. There is no direct relationship between the volumes of the hydrogen used and water produced. (There is a relationship between the number of moles.)
- (e) 2 moles of hydrogen gas will react with every 1 mole of oxygen gas to produce 2 moles of water.

1.4.4.4

Equation	$\text{N}_2(\text{g})$	+	$3\text{H}_2(\text{g})$	$\rightarrow$	$2\text{NH}_3(\text{g})$
Reacting volumes	1 vol		3 vol		2 vol
Amount at start	10 cm ³		21 cm ³		nil
Amount used or made	21/3 = 7 cm ³		21 cm ³		14 cm ³
Amount left at end	3 cm ³		0		14 cm ³

The final gas mixture will have a volume of 17 cm³ and this will include 3 cm³ nitrogen and 14 cm³ ammonia. (All of the hydrogen will have been used up.)

- 1.4.4.5**
- (a) 240.00 dm^3 ($1 \text{ dm}^3 = 1 \text{ L}$).
- (b) $240.00 \text{ dm}^3 = 240.00/22.4 = 10.71428571$
 Molar mass $\text{SO}_2 = 64.06 \text{ g}$
 $10.70950469 \text{ mol} = 10.71428571 \times 64.06 = 686.4 \text{ g}$
- 1.4.4.6**
- (a) $2\text{NO(g)} + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$
- (b) 100 cm^3 oxygen; 200 cm^3 nitrogen dioxide.
- 1.4.4.7**
- (a) $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O(g)}$
- (b) 8 dm^3
- (c) 4 dm^3
- 1.4.5.1**
- (a) Standard temperature and pressure – a temperature of 0°C (273 K) and a pressure of 101.3 kPa (1 atmosphere).
- (b) Molar gas volume is the volume occupied by 1 mole of any gas at a particular temperature and pressure. At STP the molar gas volume is $22.4 \text{ dm}^3 \text{ mol}^{-1}$. One mole of any gas will always occupy a volume of 22.4 dm^3 when the temperature is 0°C and the atmospheric pressure is 1 atmosphere (101.3 kPa).
- (c) (i) $22.4 \text{ dm}^3 = 22 \text{ dm}^3$ (with significant figures)
 (ii) $22.4 \text{ dm}^3 = 22 \text{ dm}^3$ (with significant figures)
 (iii) 22.4 dm^3

- (iv) 4.5 dm^3
 (v) 22.4 dm^3

1.4.5.2 (a) Number of moles (n) = volume/molar volume

- (b) (i) 29 mol
 (ii) 29 mol
 (iii) $9 \times 10^6 \text{ mol}$

1.4.5.3 (a) $2.80/44.01 = 0.0636 \text{ mol}$ or $6.36 \times 10^{-2} \text{ mol}$

(b) $0.0636 \times 22.4 = 1.42 \text{ dm}^3$

1.4.5.4 (a) $50 \text{ cm}^3 = 0.050 \text{ dm}^3 = 0.050/22.4 = 2.2 \times 10^{-3} \text{ mol}$

(b) Molar mass of $\text{NH}_3 = 14.01 + 3(1.01) = 17.04 \text{ g}$

(c) Mass of ammonia = no. of moles $\times$ molar mass = $2.23 \times 10^{-3} \times 17.04 = 3.8 \times 10^{-2} \text{ g}$

1.4.5.5 (a) $6\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(\text{aq}) + 6\text{O}_2(\text{g})$ (photosynthesis reaction)

(b) Volume of oxygen gas = $1.3 \text{ cm}^3 = 1.3 \times 10^{-3} \text{ dm}^3$

$$n = \text{volume/molar volume at STP} = 1.3 \times 10^{-3}/22.4 = 5.8 \times 10^{-5} \text{ moles}$$

(c) Molar mass = 32.00 g

(d) Mass = no. of moles $\times$ molar mass = $5.8035714 \times 10^{-5} \times 32.00 = 1.9 \times 10^{-3} \text{ g}$

1.4.5.6 Use all decimal places for calculations then round off for answers.

Name of gas	Mass of gas (g)	Number of moles of gas	Volume of gas at STP (dm^3)
Carbon dioxide	44.01	$44.01/44.01 = 1.00$	$1.0 \times 22.4 = 22.4$
Sulfur dioxide	44.01	$44.01/64.06 = 0.6870$	$0.687 \times 22.4 = 15.39$
Carbon dioxide	$5.00 \times 44.01 = 220.05$ $= 2.2 \times 10^2$	5.00	$5.0 \times 22.4 = 112.0$
Sulfur dioxide	$5.0 \times 64.06 = 320.30$ $= 3.2 \times 10^2$	5.00	$5.0 \times 22.4 = 112.0$
Carbon dioxide	$10.24 \times 44.01 = 450.67$	$229.48/22.4 = 10.24$	229.48

1.4.6.1 (a) Ideal gas properties:

It is made of molecules which are in constant random motion.

All collisions between the gas particles are elastic – no energy is lost.

No (or negligible) forces exist between the particles.

The volume of the gas particles is negligible

The kinetic energy of its particles is directly proportional to the absolute temperature.

An ideal gas would obey the gas law, $PV = nRT$, under all conditions of temperature and pressure.

(b) At high pressures and low temperatures gases differ most from the ideal gas model. Real gases do have forces of attraction and repulsion between their particles and their particles do have volume.

At high pressure, the particles are pushed closer together. When this happens, collisions become more frequent and the volume of particles becomes more important and eventually prevents any further compression.

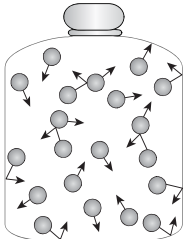
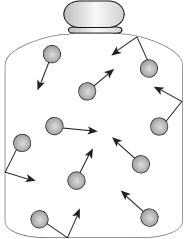
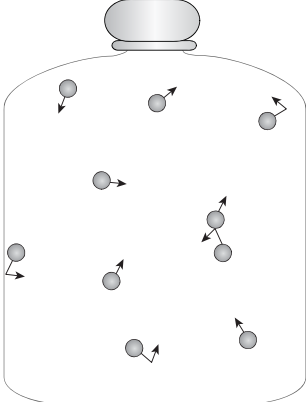
At normal temperatures gas particles are so far apart, and are moving so quickly, that the intermolecular forces between the particles have little effect. But at low temperatures, particles slow down and forces between them become significant. Thus gases do not behave like an ideal gas at low temperatures.

(c) Helium. Helium has small monatomic molecules with only very small intermolecular attractive forces whereas carbon dioxide has much larger molecules with stronger intermolecular attractive forces. Thus helium acts more like an ideal gas. The larger molecules of carbon dioxide also make carbon dioxide gas more difficult to compress and so it will liquefy at a lower pressure than gases with smaller particles and ideal gases with particles having a negligible volume.

1.4.6.2 (a) More gas in the same space means particles closer together and more collisions amongst particles as well as between particles and the container walls. Pressure is due to collisions of gas particles with the walls of the container, so pressure will increase.

(b) When the temperature is increased, the particles move faster and there are more collisions between particles and with the container walls. Thus the pressure increases.

1.4.6.3

Same container, amount of gas doubled	Same container, temperature increased	Original amount of gas added to container twice as big
 Pressure increases	 Move faster Pressure increases	 Pressure decreases

- 1.4.6.4 (a) Increase.
(b) Increase.
(c) Decrease.

1.4.6.5
$$\frac{PV_1}{T_1} = \frac{PV_2}{T_2}$$

$$\frac{105 \times 450}{300} = \frac{100 \times V_2}{273}$$

$$V_2 = 430 \text{ cm}^3$$

1.4.6.6 $T_1 = 80^\circ\text{C} = 80 + 273 = 353 \text{ K}$
 $T_2 = 30^\circ\text{C} = 30 + 273 = 303 \text{ K}$

$$\frac{PV_1}{T_1} = \frac{PV_2}{T_2}$$

$$\frac{1 \times 1.3}{353} = \frac{P_2 \times 0.5}{303}$$

$$P_2 = 2.2 \text{ atmospheres}$$

- 1.4.7.1 (a) P is pressure of the gas (kPa)
 V is volume of the gas (dm^3)
 n is number of moles of gas
 R is the universal gas constant ($8.314 \text{ J K}^{-1} \text{ mol}^{-1}$)
 T is absolute temperature (K)
- (b) $^\circ\text{C} + 273 = \text{K}$

- 1.4.7.2 (a) 0°C is the temperature at which pure water melts/freezes and 100°C is the temperature at which pure water boils at STP.
- (b) All matter is made of particles which are constantly moving (kinetic particle theory). As temperature decreases, the average velocity of particles decreases until, at a temperature of absolute zero (0 K), the particles would (theoretically) be absolutely still, they would not move at all.
- (c) Scientists have not been able to cool anything to a temperature as low as -273°C (0 K), however recently physicists took atoms to within a billionth of a degree of absolute zero by using a technique called laser cooling to trap atoms and slow their movement until they are almost motionless.

1.4.7.3

(a) $PV = nRT$

$$206 \times 40 = n \times 8.314 \times (30 + 273)$$

$$n = 3.3 \text{ moles of propane.}$$

(b) $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$

$$\frac{206 \times 40.0}{303} = \frac{101.3 \times V_2}{273}$$

$$V_2 = 73.28. \text{ At STP the propane would occupy } 73 \text{ dm}^3.$$

1.4.7.4

(a) $PV = nRT$

$$100 \times V = 11.5/32 \times 8.314 \times 373$$

$$V = 11.145. \text{ The oxygen would occupy } 11.1 \text{ dm}^3.$$

(b) Same mass, so number of moles stays the same. $T = 25 + 273 = 298 \text{ K.}$

$$100 \times V = 11.5/32 \times 8.314 \times 298$$

$$V = 8.90. \text{ The volume would be } 8.9 \text{ dm}^3.$$

1.4.7.5

(a) $PV = nRT$

$$P = \frac{nRT}{V} = \frac{250 \times 8.314 \times 300}{30} = 20\,785 \text{ kPa or } 2.08 \times 10^4 \text{ kPa}$$

(b) The actual value is lower than that calculated. The calculated value assumes that the gas behaves as an ideal gas.

Assumptions underlying the ideal gas equation that could account for the difference between the actual and calculated pressure include:

- that there are negligible forces between the molecules
- that the volume of the molecules is small enough to be negligible.

1.4.7.6

$$PV = nRT$$

$$101.3 \times 24.20 = n \times 8.314 \times 273$$

$$n = 1.080 \text{ mol}$$

$$n = \text{mass/molar mass}$$

$$1.08 = 4.321/\text{molar mass}$$

$$\text{molar mass} = 4.00$$

The gas is helium.

1.4.8.1

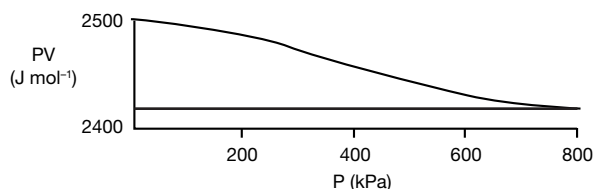
(a) For a constant mass of gas at a constant temperature, pressure is inversely proportional to its volume; or pressure is proportional to $1/V$; or $P_1 V_1 = P_2 V_2$; or $PV = \text{constant}$ (Boyle-Mariotte law).

(b) For a constant mass of gas at a constant temperature, pressure is inversely proportional to its volume; or pressure is proportional to $1/V$; or $P_1 V_1 = P_2 V_2$; or $PV = \text{constant}$. (Boyle-Mariotte law).

(c) For a constant mass of gas at a constant pressure, volume is proportional to the absolute temperature (temperature in kelvins); or $V/T = \text{constant}$; or $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ (Charles's law).

(d) For a constant mass of gas at a constant volume, pressure is proportional to the absolute temperature; or $P/T = \text{constant}$; or $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ (Gay-Lussac's law).

1.4.8.2



1.5.1.1

(a) A solution is a homogeneous mixture of two or more pure substances. The solute is evenly distributed throughout the solvent.

- (b) A solid dissolves when it is mixed into another substance (e.g. water), and 'disappears'. The liquid goes clear. (Note that clear means transparent – it may be coloured or colourless).

The liquid produced by dissolving one substance in another is a mixture (it is no longer pure), and it is called a solution. In a solution the particles of the solute are spread throughout the particles of the solvent.

A solid melts when it is heated. Its composition does not change, it is still a pure substance. The same particles are present before and after heating, the particles do not change when heated, they just move more freely.

1.5.1.2

- (a) (iii)
(b) (v)
(c) (ii)
(d) (iv)
(e) (i)

1.5.1.3

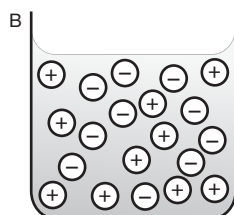
- (a) Saturated solution.
(b) Dilute solution.
(c) Concentrated solution.

1.5.1.4

Various, e.g.

Solvent	Substance that could dissolve in the solvent
Water	Salt (sodium chloride), sodium hydrogen carbonate, some dirt on hands and clothes, carbon dioxide in soft drinks.
Nail polish remover (ethyl acetate)	Nail polish.
Methylated spirits (mostly ethanol)	Some inks, oil and grease.
Eucalyptus oil	Glue on bottles (from labels) and grease on clothes.
Turpentine	House paint.
Petrol	Oil in lawnmower fuel.

1.5.1.5



1.5.1.6

48 milligrams per millilitre = 48×10^{-3} g per mL = 48 g dm^{-3}

1.5.1.7

- (a) $[\text{NaCl(aq)}] = 10 \text{ g dm}^{-3}$
(b) $22.99 + 35.45 = 58.44 \text{ g}$
(c) $10 \text{ g} = 10/58.44 \text{ mol} = 0.17111 \text{ mol}$
Concentration = 0.17 mol dm^{-3}
(d) Sodium chloride is an ionic salt; you cannot calculate the mass of a molecule as it does not exist as molecules.

1.5.1.8

- (a) 1.2 g
(b) 39.998 g
(c) $0.060 \text{ mol dm}^{-3}$

1.5.1.9

- (a) 52.5 g
(b) 74.44 g
(c) $0.564 \text{ mol dm}^{-3}$

1.5.1.10

- (a) 3 mol
(b) 0.50 mol
(c) 0.15 mol

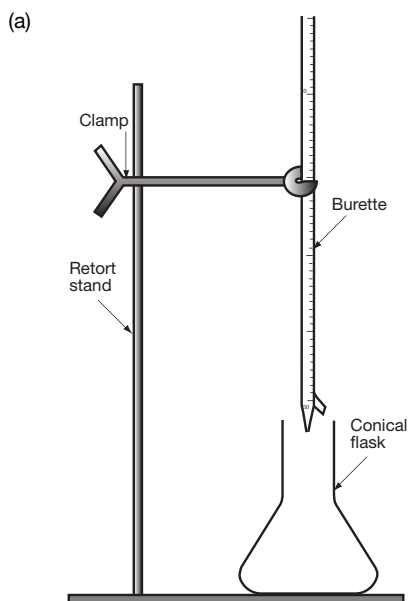
- 1.5.1.11** (a) 0.02 mol
(b) 95.21 g
(c) 1.9 g
- 1.5.1.12** (a) 1 mol $\text{Na}_2\text{CO}_3 = 105.99 \text{ g}$
Mass needed to make 1 dm^3 of 0.50 mol dm^{-3} solution = 52.995 g
For 500 cm^3 use $52.995 \times 500/1000 = 26.4975 \text{ g}$
Weigh accurately 26.4975 g Na_2CO_3 , dissolve in water and make up to 600 cm^3 accurately, e.g. using volumetric flask.
(b) $100 \times 0.05 = 2.00 \times V$
 $V = 2.5 \text{ cm}^3$
Measure accurately 2.50 cm^3 of 2.00 mol dm^{-3} hydrochloric acid.
Add water until the volume of the solution is 100 cm^3 .
- 1.5.2.1** (a) $\text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$
(b) Molar mass of Mg = 24.31 g
Number of moles = $1.73/24.31 = 0.0712 \text{ mol}$
(c) 0.0712 mol hydrogen
(d) Molar mass of hydrogen gas = 2.016 g
Mass of hydrogen = $2.016 \times 0.071164129 = 0.143 \text{ g}$
(e) Molar volume at STP = 22.4 dm^3
Volume of hydrogen = $22.4 \times 0.071164129 = 1.59 \text{ dm}^3$
- 1.5.2.2** (a) $\text{H}_2\text{SO}_4\text{(aq)} + 2\text{NaOH(aq)} \rightarrow 2\text{H}_2\text{O(l)} + \text{Na}_2\text{SO}_4\text{(aq)}$
(b) $n = cV = 0.20 \times 50/1000 = 0.01 \text{ mol H}_2\text{SO}_4$
(c) 0.02 mol NaOH
(d) $n = cV$
 $c = n/V = 0.02/(25 \times 10^{-3}) = 0.80 \text{ mol dm}^{-3}$
(e) Add an indicator which will change colour at the relevant pH.
- 1.5.2.3** (a) Concentrated diluted acid = 0.60 mol dm^{-3}
Concentrated original acid = 3.0 mol dm^{-3}
(b) The first titration would be rough. Its role would be to give them a quick idea of approximately how much sodium hydroxide would be needed for complete reaction. The next three titrations could then be accurate and should be within 0.1 cm^3 of each other. Discarding the rough result and averaging the three accurate results increases the accuracy.
- 1.5.2.4** (a) Volumetric analysis – finding accurately the composition of a solution by measuring. Volumetric means that volumes are measured in a technique such as a titration.
(b) Standard solution – the reactant solution, in a titration, whose concentration is accurately known.
(c) Equivalence point – the point when the reaction is complete. At this point, the reactants and products are in the stoichiometric ratio shown by the balanced equation.
(d) End point – when the indicator changes colour. If the indicator is chosen correctly, the end and equivalence points will coincide, the indicator changing colour when the reaction is complete.
- 1.5.2.5** (a) A primary standard is a solution that is made by dissolving an accurately measured mass of a solute in a small amount of the solvent and making the volume up to a measured volume. A secondary standard is a solution whose concentration is determined by titration against a primary standard.
(b) A primary standard:
- can be obtained in a pure form (Common acids such as HCl and H_2SO_4 would not be suitable as their concentration varies from batch to batch.)
 - has a known chemical formula
 - is stable and does not change when exposed to air (sodium hydroxide would not be suitable as it absorbs water from air and reacts with carbon dioxide in the air)
 - is soluble.

- (c) The most commonly used primary standard is anhydrous sodium carbonate (Na_2CO_3). This makes a basic solution, suitable for analysing acids. A suitable chemical to use as an acidic primary standard would be hydrated oxalic acid ($\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$).
- (d) Sodium hydroxide pellets are deliquescent – they absorb water from the air; also atmospheric carbon dioxide will dissolve in sodium hydroxide. Thus you cannot prepare the pure solution of NaOH with a definite concentration. A definite concentration is required for a primary standard, so NaOH is not suitable to use for this purpose.

1.5.2.6

Acid	Base	Suitable indicator	Justification
HCl	KOH	Bromothymol blue	Strong acid/strong base so equivalence point approximately = 7. Bromothymol blue changes colour around pH = 7.
H_2SO_4	NH_4OH	Methyl orange	Strong acid/weak base so equivalence point < 7. Methyl orange changes colour at pH < 7.
H_2CO_3	NaOH	Phenolphthalein	Weak acid/strong base so equivalence point > 7. Phenolphthalein changes colour at pH > 7.
CH_3COOH	$\text{Ca}(\text{OH})_2$	Phenolphthalein	Weak acid/strong base so equivalence point > 7. Phenolphthalein changes colour at pH > 7.

1.5.2.7



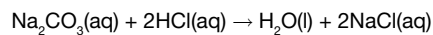
- (b) Volumetric flask.
- (c) Wash bottle (distilled water bottle).
- (d) Pipette.
- (e) Pipette filler.

1.5.2.8

Factor	Pipette	Burette
Function	Accurately measures a fixed volume of solution to provide a fixed number of moles of one reactant.	Allows you to measure the exact volume, of a reactant, needed to reach the equivalence point.
Procedure for washing before a titration	Pipette must be washed with distilled water and then with the solution to be used in it.	Burette should be washed with distilled water and then with the solution to be used in it.
Name of measured volume	An aliquot.	A titre.

- 1.5.2.9**
- (a) Rinsing the flask with the solution would produce inaccurate results. The film of solution left on the inside of the conical flask would increase the number of moles of that chemical in the flask.
- Rinsing with water would not affect accuracy. A film of water left on the inside of the flask would not matter, because this will not change the number of moles of the chemicals being added to the flask.
- (b) The pipette is calibrated to allow the required volume to be released without the addition of the drop(s) left in the bottom.
- (c) Various, e.g.
- Avoid parallax error when reading the volume in the burette and pipette by making sure your eyes are at the same level as the meniscus you are reading.
- When making up a standard solution, use a wash bottle with distilled water to wash any remaining solid into the beaker and from the beaker into the volumetric flask.
- Always average the three accurate titration readings.
- Choose the indicator carefully so it changes colour (end point) as close as possible to the equivalence point of the reaction.
- 1.5.2.10**
- (a) Strong base and strong acid.
- (b) Strong base and weak acid.
- (c) Weak base and strong acid.
- (d) Weak base and weak acid.
- 1.5.2.11**
- The graph starts off fairly level, sloping down only slightly – this indicates a slow drop in pH while the first 24 mL of HCl is added and the sodium hydroxide is being neutralised.
- During the addition of the next 1.5 mL there is a sudden drop in pH from 11.31 to 3.00. This shows that the reaction is complete – the equivalence point is reached at the centre of this drop in the graph. The middle of this sharp drop is at pH = 7, indicating a reaction between a strong base and a strong acid and the formation of a neutral salt (sodium chloride).
- The graph then levels out, curving down slightly as excess acid is added and the acidity increases slightly (lower pH).
- 1.5.2.12**
- Potassium permanganate solution contains the permanganate ion (MnO_4^-). This is a strong oxidising agent. In acid solution, the reaction is:
- $$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$$
- The MnO_4^- ion is purple and the Mn^{2+} ion is almost colourless.
- If the permanganate (in a burette) is titrated against a reducing agent (in a conical flask) with colourless ions, then as soon as the reducing agent has all reacted, the next drop of permanganate will produce a purple colour in the flask. It changes colour itself at the end point (when the purple permanganate ion is converted to the colourless manganese ion), so no other indicator is needed.
- 1.5.2.13**
- (a) $\text{NaOH}(\text{aq}) + \text{HCl}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
10.0 cm³
- (b) $2\text{NaOH}(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
20.0 cm³
- 1.5.2.14**
- (a) (i) 105.99 g
(ii) 0.03 moles
(iii) 0.12 mol dm⁻³
(full value for further calculations is 0.122265308)
- (b) (i) 3.07 moles
(ii) 6.13 moles
(iii) 0.5 mol dm⁻³
- 1.5.2.15**
- (a) Number of moles of sodium carbonate (Na_2CO_3)
- $$= \frac{\text{mass}}{\text{molar mass}} = \frac{12.9}{2 \times 22.99 + 12.01 + 3 \times 16.0}$$
- $$= 0.1217 \text{ moles in } 250 \text{ cm}^3$$
- $$\text{Concentration (moles per dm}^3\text{)} = 0.1217 \times \frac{1000}{250}$$
- $$= 0.487 \text{ moles per dm}^3$$
- (Note: It is good to round off for this answer, but always use the full number in the calculator for any further calculations.)

(b) Write the equation:



Calculate number of moles:

$$\begin{aligned}\text{Number of moles of Na}_2\text{CO}_3 &= \text{concentration} \times \text{volume} \\ &= 0.48683838 \times 0.025 \\ &= 0.01217 \text{ moles}\end{aligned}$$

$$\begin{aligned}\text{Number of moles of HCl} &= \text{concentration} \times \text{volume} \\ &= \text{concentration} \times 0.0249\end{aligned}$$

Use the equation:

1 mol Na_2CO_3 neutralises 2 mol HCl, so

0.01217 mol Na_2CO_3 neutralises 2×0.01217 mol HCl.

Calculate concentration:

$$\text{Number of moles of HCl} = \text{concentration} \times 0.0249,$$

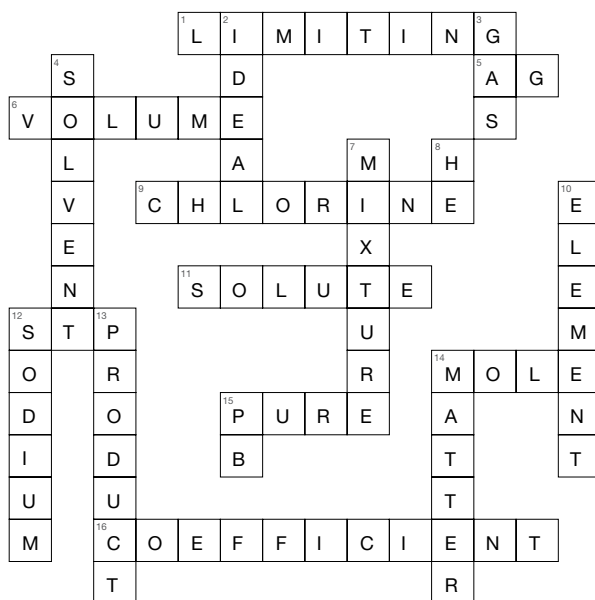
$$\text{so } 2 \times 0.01217 = \text{concentration} \times 0.0249$$

$$\text{Concentration of HCl} = 0.978 \text{ moles per dm}^3$$

- 1.5.2.16**
- (a) $\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{Fe}^{2+}(\text{aq}) \rightarrow 5\text{Fe}^{3+}(\text{aq}) + \text{Mn}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- (b) (i) 2.77×10^{-3} mol MnO_4^-
(ii) 1.39×10^{-2} mol Fe^{2+}
- (c) 7.76×10^{-1} grams iron
- (d) 64.67%

Just for fun

- Reactions can be stoichiometric.
- Empirical.
 - Synthesis.
 - Compounds.
 - Spectator.
 - Subscript.
- Crossword.



[illegible]